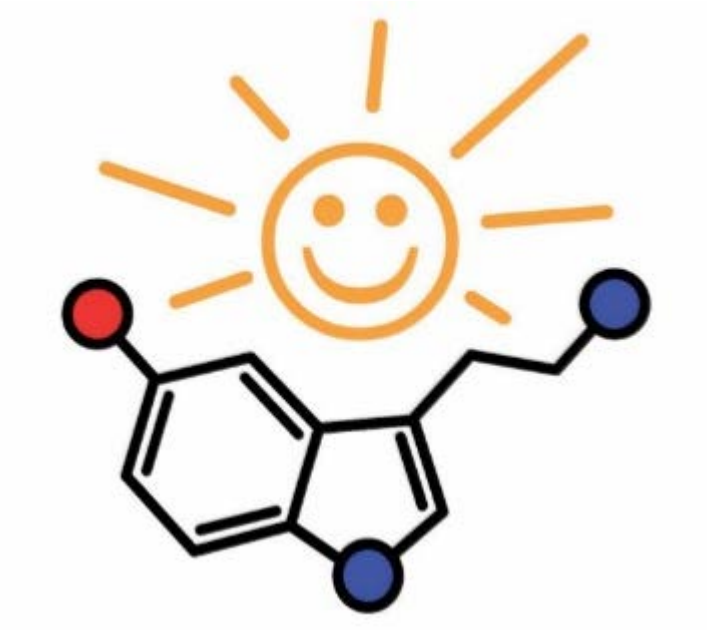
Student Name: _____

Class period: _____

RRHS AP Chemistry



Unit 10:

Atomic Structure & Periodicity

Unit 11: Bonding & IMFs

Table of Contents

Class Info

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B Day: Text **@APLestikB** to **81010**

Mastering Chemistry (Pearson textbook/homework)

https://www.pearsonmylabandmastering.com/northamerica/masteringchemistry/

You will need to get your course ID and access code from your teacher!

AP Chem: Effective Study Skills Tips and Tricks!

Study smarter, not harder. ☺

What to Do	What **NOT** to Do
Be **ACTIVE** in while learning/studying: • Close your booklet and try problems on your own with just a periodic table and formula chart! Only check your answer/work when you've finished, or you can't go any farther. • Use flashcards (physical or digital) • Struggle with challenging problems and keep trying, even if you're stuck initially (or convinced you're doing it wrong)	Be passive while learning/studying: • Re-read over your booklet and practice problems you've already completed • Ask your friend or look up the answer if you don't immediately know how to do the problem
Focus when studying • Decrease distractions while studying; don't read texts, check social media, or watch Netflix while studying. Put your phone out of sight/hearing.	Multitask • Study while checking/writing texts, checking social media, and/or watching Netflix. • Keep your computer or tv on in the background
Use **Intensity** when studying • You control the effort that you apply in your work! 30 minutes of high focus, high intensity study can be better than 2 hours of unfocused, low energy multi-tasking.	Low intensity/low effort • Look over problems and try them "in your head" but then just look up the answer • Use flashcards but don't try to recall the info on the other side before looking at the answer
Space out studying over time • Study a little bit of chemistry most days • Start long-term homework (like Mastering Chem or lab reports) the day they're assigned, and work a little bit every day or two • Less is more! Spaced practice studying is more effective than LONG hours of "studying" with multitasking and little focus.	Cram • Only study for quizzes/tests the night before • Start Mastering Chem or your lab report only 1-2 days before it's due • Study for many hours at a time all at once
Interleave your Studying • Study more than just one type of problem; mix it up and jump between different concepts • Review and practice old units while studying (especially important since AP Chem assessments are cumulative!)	One Concept Studying • Study only one type of problem, and practice those problems over and over • Don't review older content or units while studying

Test Yourself!
• The best way to prepare for a test is to take a test!
  ○ Time yourself while trying practice problems
  ○ Access only the AP Periodic Table and Formula Chart when practicing problems
	Open Notes Practice
	• Use your notes, friends, and/or the internet while trying practice problems
	• Give yourself unlimited time for each problem

4th Marking Period: March – May 2023

Sunday	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday
26	**27 A**	**28 B**	**1 Mar. A** **SAT (11th only)**	**2 B**	**3 A**	4
5	**6 B**	**7 A**	**8 B**	**9 A**	**10 B** **MP3 ENDS**	11
12	**13** **SPRING**	**14** **BREAK**	**15** **SPRING**	**16** **BREAK**	**17** ☺	18
19	**20 A**	**21 B**	**22 A**	**23 B**	**24 A**	25
26	**27 B**	**28 A**	**29 B**	**30 A**	**31 B**	1 Apr.
2	**3 A**	**4 B**	**5 A**	**6 B**	**7** Student & Staff Holiday	8

Sunday	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday
9 Apr	**10** **A**	**11** **B**	**12** **A**	**13** **B**	**14** **A**	15
16	**17** **B**	**18** **A** **STAAR English I Exam**	**19** **B**	**20** **A** **STAAR English II Exam**	**21** Student Holiday/ Staff Development	22
23	**24** **B**	**25** **A** STAAR Biology EOC	**26** **B**	**27** **A** STAAR US Hist EOC	**28** **B**	29
30	**1** May **A** AP Chem Exam! ;D	**2** **B** STAAR Alg I EOC	**3** **A**	**4** **B**	**5** **A**	6
7	**8** **B**	**9** **A**	**10** **B**	**11** **A** WHAP Exam: some 10th	**12** **B**	13
14	**15** **A**	**16** **B**	**17** **A** Seniors: 3rd & 4th Exams in class	**18** **B** Seniors: 7th and 8th Exams in class	**19** **C** Last day for obligations **MP4 ENDS**	20
21	**22** **Exams 2, 1**	**23** **Exams 6, 5**	**24** **Exams 3, 4**	**25** **Exams 7, 8**	**26**	27

AP® CHEMISTRY EQUATIONS AND CONSTANTS

Throughout the exam the following symbols have the definitions specified unless otherwise noted.

L, mL	=	liter(s), milliliter(s)	mm Hg	=	millimeters of mercury
g	=	gram(s)	J, kJ	=	joule(s), kilojoule(s)
nm	=	nanometer(s)	V	=	volt(s)
atm	=	atmosphere(s)	mol	=	mole(s)

ATOMIC STRUCTURE

$E = h\nu$

$c = \lambda\nu$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s
Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number = 6.022×10^{23} mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$, where $a\,A + b\,B \rightleftarrows c\,C + d\,D$

$K_p = \frac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$

$K_a = \frac{[H^+][A^-]}{[HA]}$

$K_b = \frac{[OH^-][HB^+]}{[B]}$

$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C

$= K_a \times K_b$

$pH = -\log[H^+]$, $pOH = -\log[OH^-]$

$14 = pH + pOH$

$pH = pK_a + \log\frac{[A^-]}{[HA]}$

$pK_a = -\log K_a$, $pK_b = -\log K_b$

Equilibrium Constants

K_c (molar concentrations)
K_p (gas pressures)
K_a (weak acid)
K_b (weak base)
K_w (water)

KINETICS

$\ln[A]_t - \ln[A]_0 = -kt$

$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$

$t_{1/2} = \frac{0.693}{k}$

k = rate constant
t = time
$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

$$n = \frac{m}{M}$$

$$\text{K} = {}^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

Molarity, M = moles of solute per liter of solution

$$A = abc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
M = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
a = molar absorptivity
b = path length
c = concentration

Gas constant, R = 8.314 J mol^{-1} K^{-1}
= 0.08206 L atm mol^{-1} K^{-1}
= 62.36 L torr mol^{-1} K^{-1}

1 atm = 760 mm Hg = 760 torr

STP = 273.15 K and 1.0 atm

Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$
$$= -RT \ln K$$
$$= -nFE^\circ$$

$$I = \frac{q}{t}$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

Unit 10 Objectives: Atomic Structure & Periodicity

BIG IDEA 1 - The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identity in chemical reactions.

- Enduring Understanding 1.A: All matter is made of atoms. There are a limited number of types of atoms; theses are elements.
- Enduring Understanding1.B: The atoms of each element have unique structures arising from interactions between electrons and nuclei.
- Enduring Understanding 1.C: Elements display periodicity in their properties when the elements are organized according to increasing atomic number. This periodicity can be explained by the regular variations that occur in the electronic structures of atoms. Periodicity is a useful principle for understanding properties and predicting trends in periods. Its modern-day uses range from examining the composition of materials to generating ideas for designing new materials.
- Enduring Understanding 1.D: Atoms are so small that they are difficult to study directly; atomic models are constructed to explain experimental data on collections of atoms.
- Enduring Understanding 1.E: Atoms are conserved in physical and chemical processes.

BIG IDEA 5 - The laws of thermodynamics describe the essential role of energy and explain and predict the direction of changes in matter.

- Enduring Understanding 5.B: Energy is neither created nor destroyed, but only transformed from one form to another.
- Enduring Understanding 5.D: Electrostatic forces exist between molecules as well as between atoms or ions, and breaking the resultant intermolecular interactions requires energy.

Coulomb's Law: fundamental relationship between electrostatic _____ and _____.

- It applies to charged particles, magnets, gravitation, etc.
- In chemistry, we are most interested in _____ of attraction or repulsion between _____ particles

Coulomb's Law: $E \propto \frac{Q_1 Q_2}{r}$

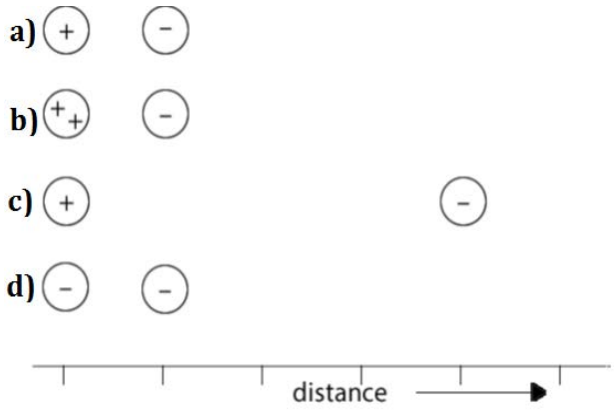
E= energy of attraction or repulsion between particles
Q_1 = charge of first particle
Q_2 = charge of second particle
r = distance between charged particles

In short:

- Energy of attraction/repulsion _____ as the magnitudes (sizes) of the charges _____
- Energy of attraction/repulsion _____ as the distance between the charges _____

Thought question: will an electron be more attracted to the nucleus of a hydrogen atom or a helium atom, and why?

Coulomb's Law Practice



1. Which set of particles shown the left will experience:
   a. the greatest attraction to each other? _____
   b. the greatest repulsion from each other? _____
2. The particles in (a) will experience _____ attraction to each other than the particles in (c) because:
   a. greater, the distance between them is less.
   b. smaller, the distance between them is less.
   c. the same, distance is irrelevant to force of attraction.
3. The particles shown above in (a) will experience _____ attraction to each other than the particles in (b) because:
   a. greater, the nucleus has one proton instead of two.
   b. smaller, the nucleus has one proton instead of two.
   c. the same, charge is irrelevant to force of attraction.
4. An electron in the lowest energy level would be most attracted to the nucleus of which element?
   a. lithium
   b. sodium
   c. potassium
   d. rubidium
5. A proton would be <u>least</u> repulsed by the nucleus of which element?
   a. helium
   b. neon
   c. argon
   d. krypton

Atomic Structure

The atom can be divided into two regions:

- Nucleus: a very small region near the center of an atom that has a positive charge. It contains:
  - Protons: charge: ____, mass: _____; the number of protons is known as the _____ number and distinguishes one element from another.
  - Neutrons: charge: ____, mass: _____

***Note**: the unit of mass for atomic particles is the atomic mass unit (_____)

→ 1 amu = one-twelfth the mass of a carbon atom containing six protons and six neutrons.

- Electron cloud: a very large region surrounding the nucleus that is negatively charged. It contains:
  - Electrons: charge: ____, mass: _____; atoms that are electrically neutral (they have no charge) must contain the _____ number of protons and electrons.
  - It consists mostly of empty space.
  - Electrons exist in complex regions of space known as _____, which are organized in various energy levels.

Isotopes

What are **isotopes**? Atoms of the _____ element, but different _____

This means the number of _____ is the same, and the number of _____ is different.

➔ **Mass of an isotope = # protons + # neutrons**

Two ways to write isotopes:

Type	**hyphen-notation**	**VS**	**isotope notation/ nuclide symbol**
Definition	name-mass		$^{\text{mass \#}}_{\text{atomic \#}}\text{Symbol}$
Example	carbon-12		$^{12}_{6}\text{C}$

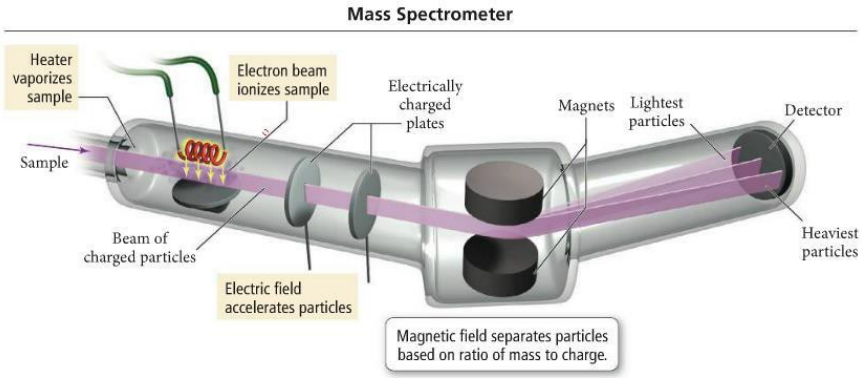
Let's Practice with Tasty Isotopes! Complete the chart below.

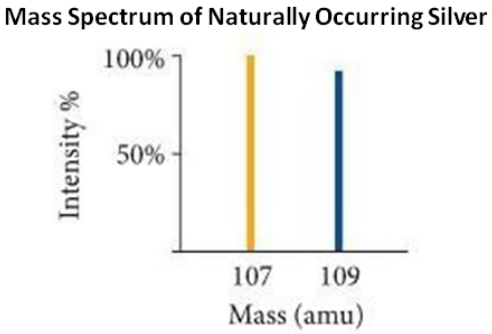
Hyphen notation	Isotope Notation	Mass #	Atomic #	Protons	Neutrons	Electrons
	$^{13}_{6}\text{C}$					
phosphorus–32						
		66		30		

Mass Spectrometry and Average Atomic Mass

The masses of elements and their percent abundances of isotopes of elements are measured using _____

spectrometry: a technique that separates particles according to their mass, producing a mass _____.





a. The <u>position</u> (location) of each _____on the x-axis indicates the _____ of the isotope.

b. The <u>intensity</u> (indicated by the_____of the peak) indicates the <u>relative abundance</u> (how common that isotope is in _____).

<u>Average Atomic Mass</u>: the weighted average mass of an element's isotopes and is the mass found on the periodic table.

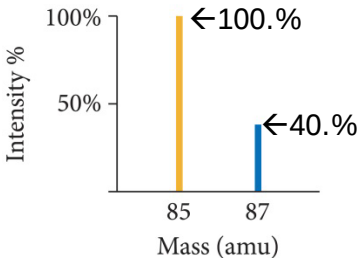
$$Average\ atomic\ mass = mass_1\left(\frac{\%\ Abundance}{100}\right) + mass_2\left(\frac{\%\ Abundance}{100}\right) + \cdots$$

- The average atomic mass will be _____ the mass of the largest and the mass of the smallest isotope.
- The average atomic mass will generally be _____ to the most abundant isotope.
- <u>Note</u>: It is important to understand that the masses of a proton and neutron are ***approximately*** 1 amu, but the actual mass of each isotope is _____ a whole number (mmm, nuclear binding energy). When specific, non-whole number masses are provided for each isotope, use the specific masses!

<u>Free Response Practice</u>:

1. Consider the mass spectrum shown below.

a. Determine the average atomic mass of the element using the mass spec shown below.



b. Which element is most likely shown here? _____

c. How did you identify the element?

2. There are two stable isotopes of calcium: Ca-40 (39.96) and Ca-46 (45.95).

a. Using the average atomic mass of calcium from the periodic table, calculate the % abundance of each isotope of calcium.

b. How many atoms of Ca-46 would be found in a 25.0 g sample of naturally occurring calcium?

3. The mass reported on the periodic table for chlorine is 35.45 amu. Why, when a sample of chlorine is examined with a mass spectrometer, is there no peak in the spectrum with a mass of 35.45 amu?

4. Silver consists of two stable isotopes, one with a mass of 106.905 and an abundance of 51.84%.

a. What is the abundance and mass of the other isotope?

b. How many silver-107 atoms are present in a 2.00 gram sample of pure silver?

5. Sulfur has four isotopes: ^{32}S, ^{33}S, ^{34}S, and ^{36}S.

a. Which isotope would you predict to be the most abundant and why? (No math needed! 😊)

b. What is the difference between these four isotopes?

Multiple Choice Practice: Average Atomic Mass and Atomic Structure, Yum!

6. The average atomic mass of naturally occurring neon is 20.18 amu. There are two common isotopes of naturally occurring neon as indicated in the table below. Using the information provided, calculate the percent abundance of each isotope.

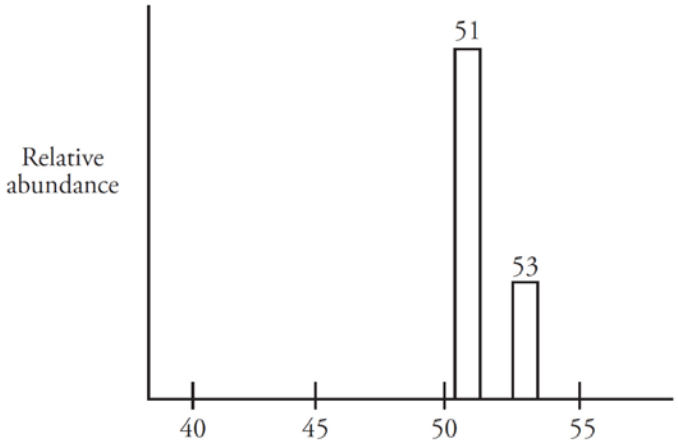
Isotope	Mass (amu)
Ne-20	19.99
Ne-22	21.99

a. 90.5% ^{20}Ne and 9.5% ^{22}Ne

b. 98.2% ^{20}Ne and 1.2% ^{22}Ne

c. 10.5% ^{20}Ne and 89.5% ^{22}Ne

d. 56.4% ^{20}Ne and 43.6% ^{22}Ne



7. The above mass spectrum is for the hypochlorite ion, ClO^-. Oxygen has only one stable isotope, which has a mass of 16 amu. Using the spectrum, calculate the average mass of a hypochlorite ion.

c. 52.5 amu b. 52.0 amu c. 51.5 amu d. 51.1 amu

8. Which of the following represents a pair of isotopes?

a. Atom #1: atomic # = 6, mass # = 14; Atom #2: atomic # = 7, mass # = 14

b. Atom #1: atomic # = 6, mass # = 7; Atom #2: atomic # = 14, mass # = 14

c. Atom #1: atomic # = 6, mass # = 14; Atom #2: atomic # = 14, mass # = 28

d. Atom #1: atomic # = 7, mass # = 13; Atom #2: atomic # = 7, mass # = 14

9. The table below shows the atomic mass and natural abundance of the two naturally occurring isotopes of lithium.

Naturally Occurring Isotopes of Lithium

Isotope	Atomic Mass (u)	Natural Abundance (%)
Li-6	6.015	7.6
Li-7	7.016	92.4

a. (7.6)(6.015 amu) + (92.4)(7.016 amu)

b. $\frac{(7.6)(6.015\ amu)+(92.4)(7.016\ amu)}{2}$

c. (0.076)(6.015 amu) + (0.924)(7.016 amu)

d. $\frac{(0.076)(6.015\ amu)+(0.924)(7.016\ amu)}{2}$

Use the table shown below to answer questions #10-11.

Data Table of Isotopes			
Set	Number of protons	Number of neutrons	Number of electrons
1	16	16	16
2	15	16	15
3	16	15	16
4	15	15	16

10. Which set of values is correct for a neutral phosphorus atom?

a. 1 b. 2 c. 3 d. 4

11. Which set(s) correctly identify an isotope of sulfur?

a. 2 only b. 1 and 3 c. 1 and 4 d. 1, 2 and 3

12. If the abundance of ^{6}Li (6.015 amu) is 7.50% and the abundance of ^{7}Li (7.016 amu) is 92.5%, what is the average atomic mass?

a. 6.08 amu b. 6.75 amu c. 6.93 amu d. 6.98 amu

Orbitals and Electron Configuration

Electron Orbitals: three-dimensional regions around the nucleus which indicate the _____ location of the electron in the electron cloud.

- Erwin Schrödinger used the hypothesis that electrons have a dual wave-particle nature to develop wave equations (or wave functions) to describe electrons. ($H\Psi = E\Psi$)
- Each individual orbital is located in a specific main energy level *and* sublevel
  - **Main energy levels or shells** (_____, the principle quantum number)
    - the general amount of _____ and _____ from the nucleus a given electron in an orbital possesses.
    - Each _____, or row, on the periodic table indicates a main energy level.
    - Indicated by numbers
  - **Energy sublevels or subshells** (*ℓ*, the angular momentum quantum number):
    - the different _____ of orbitals that exist within the same main energy level
    - Indicated by letters
- A maximum of _____ electrons can fit in a single orbital.

Electron configuration: the arrangement of electrons in an atom

- Every single atom and ion has a specific electron configuration that tells you the exact _____ and _____ of each electron it contains!

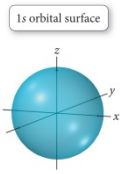
Energy Sublevels

Orbital/sublevel name (*lowercase)*	Shape	How many orbitals in that sublevel? (per main energy level)	max # of e^-?

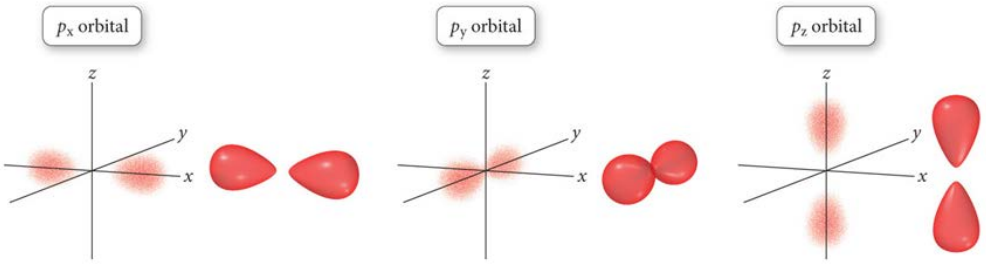
How to remember the order of the energy sublevels:

So what do these pesky orbitals actually look like?

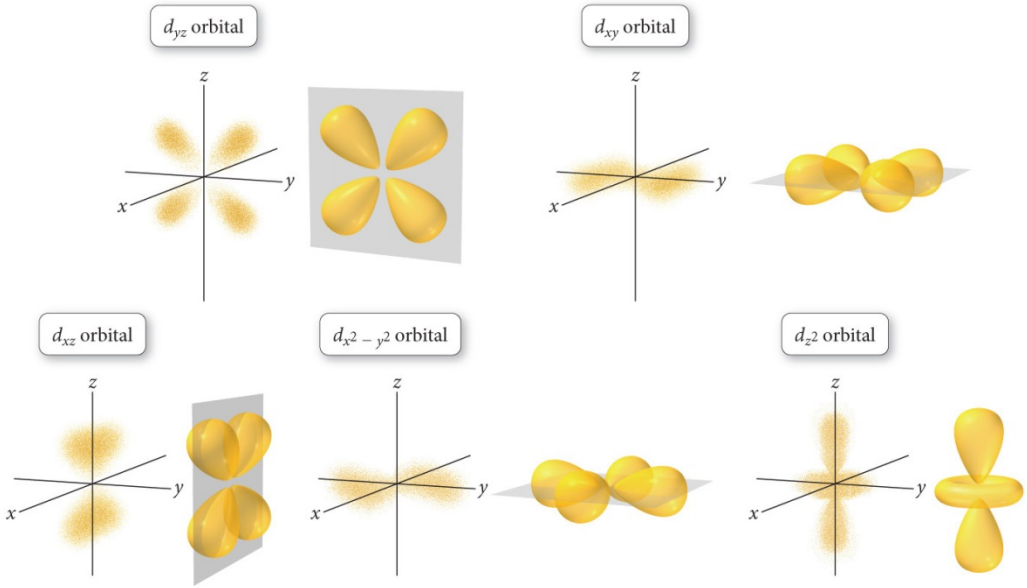
One s orbital



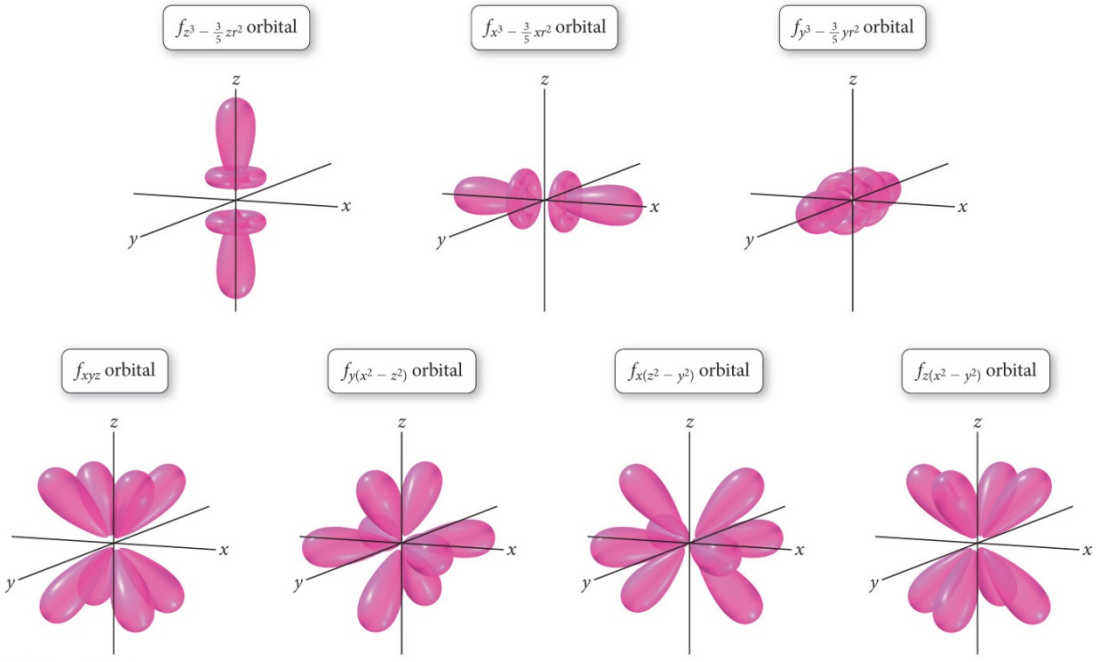
Three p orbitals

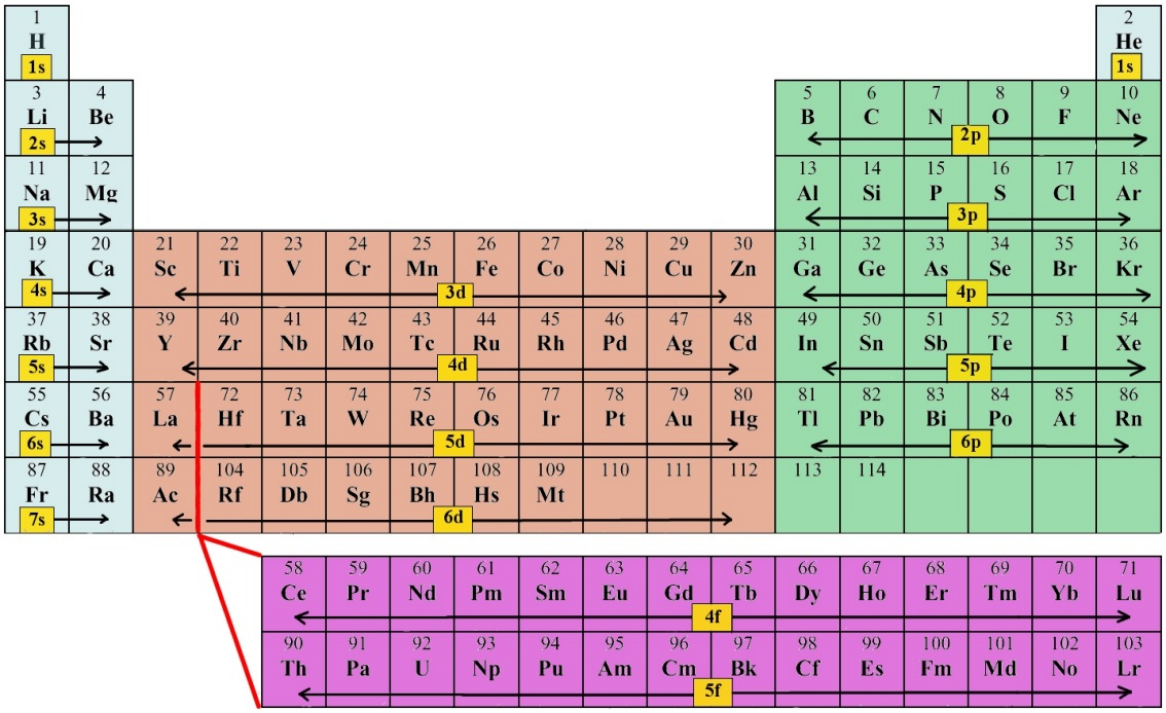


Five d orbitals

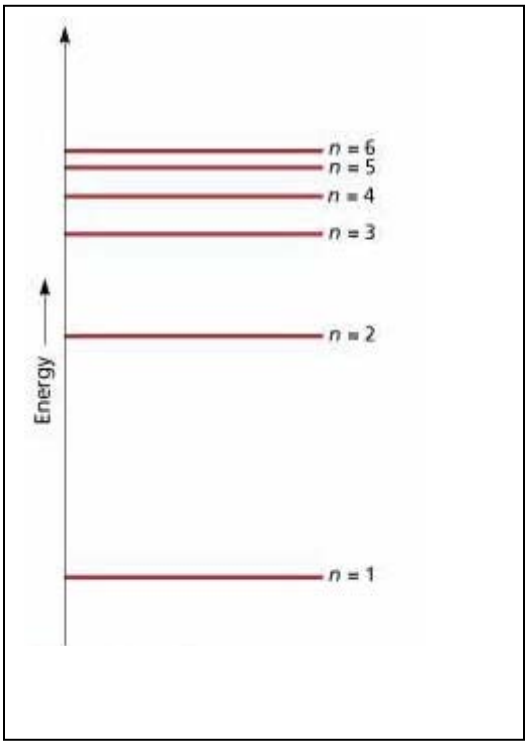


Seven f orbitals





Note: as one moves further from the nucleus, the energy levels get closer together and there is a smaller difference in energy between the levels.



Electronic Transitions

a. Absorption: electron jumps UP to a higher energy level

b. Emission: electron drops DOWN to a lower energy level

Directions: Match each of the following electronic transitions with the most likely energy change.

Transition	Energy Change
1. n = 1 to n = 4 _____	A. red light emission
2. n = 5 to n = 4 _____	B. red light absorption
3. n = 4 to n = 1 _____	C. UV light absorption
4. n = 4 to n = 5 _____	D. UV light emission

Rules Governing How Electrons Fill Orbitals in the Ground State (lowest energy state)

The **Heisenberg Uncertainty Principle** states that it is impossible to determine simultaneously both the position and the velocity of an electron or any other particle. However, we are able to determine the *probable* location of an electron and determine how electron orbitals are filled, using the following rules:

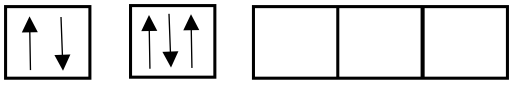
1. **Aufbau Principle**: (German for "_____") states that, in the ground state (lowest energy state), an electron will occupy the lowest-energy orbital that can receive it.

2. **Pauli Exclusion Principle:** no two electrons can fit into the same orbital with the same spin (i.e. no two electrons can have the same set of four quantum numbers!) For the new AP CHEM exam, you do NOT have to answer questions about quantum numbers. So, basically, recognize that no two electrons can exist in the exact same orbital having the exact same spin.

3. **Hund's Rule:** (the _____ rule) states that orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and all electrons in singly occupied orbitals must have the same spin. "Spin" is designated by an arrow: **spin up** = _____ or _____; **spin down** = _____ or _____

Practice! For each **ground state** orbital configuration shown below, identify which (if any) of the rules govern electron-filling have been violated (Pauli exclusion principle, Aufbau principle, and/or Hund's rule).

Violates?

1. 

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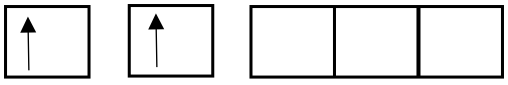
2. 

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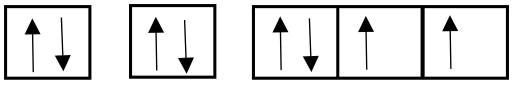
3. 

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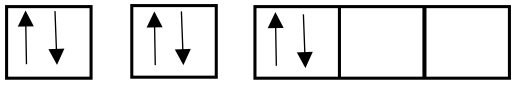
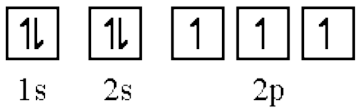
4. 

_____

Electron Configurations: The arrangement of electrons in an atom

A. Orbital Diagram: uses arrows to represent electrons. N: 

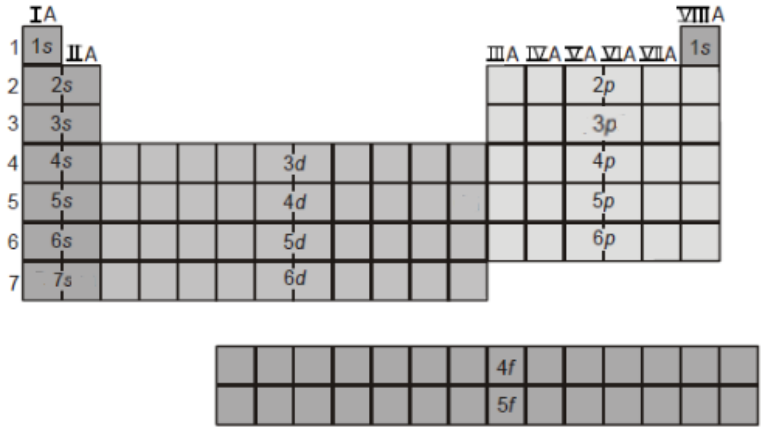
B. Complete Electron Configuration: Standard electron configurations eliminate the lines and arrows of orbital notation. The number of electrons in a sublevel is shown by adding a superscript.

Nitrogen: $1s^2 2s^2 2p^3$

C. Noble Gas Electron Configurations: A noble gas configuration is an electron configuration that utilizes a noble gas which has its valence level fully occupied. Noble gas configurations are often used to help shorten the electron configurations of those elements that contain large numbers of electrons.

Nitrogen: $[He]2s^2 2p^3$

Adding Electrons: The periodic table holds the answer for the order in which electrons _____ orbitals. Always start at hydrogen and then move through the periods (rows) until you arrive at the desired element.



The Order in Which Electrons FILL Energy Levels

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^{10}$

Order in which electrons are removed: Backwards! *EXCEPT "d" orbitals*

➔ Remove p and s electrons BEFORE d electrons (on the same energy level)

Example #1: Iron

Fe atom = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ [Ar] 4s  3d

When iron forms a cation, it FIRST loses its valence electrons FROM THE 4s SUBLEVEL:

Fe^{2+} cation = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$ [Ar] 4s _____ 3d 

It can THEN lose 3*d* electrons:

Fe^{3+} cation = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$ [Ar] 4s _____ 3d 

Example #2: Gallium

Ga atom = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$

Ga^{3+} cation = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$

<u>Isoelectronic</u> atoms and ions: the "iso" in "isoelectronic" means "_____", so isoelectronic atoms and ions have the same number of _____.

Example:

- Write the electron configuration for S^{2-}:

- What noble gas is isoelectronic to S^{2-}?

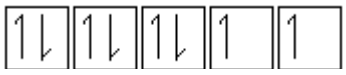
- List other atoms or ions that are isoelectronic to the S^{2-} ion:

Yum, atoms! Let's practice.

Part I: The counting of electrons.

Si	Orbital Notation	
	Electron Configuration	Noble Gas Configuration
As	Orbital Notation	
	Electron Configuration	Noble Gas Configuration
Cl^-	Orbital Notation	
	Electron Configuration	Noble Gas Configuration
In^{3+}	Orbital Notation	
	Electron Configuration	Noble Gas Configuration

Part II: Identify the atoms by examining their arrangement of electrons.

Orbital Diagram or Electron Configuration	# of Total Electrons	# of Valence Electrons	Element
			
[Ne] $3s^2 3p^2$			
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$			
[Ar] $4s^2 3d^{10} 4p^5$			
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^3$			

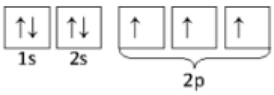
Part III: Matching!

A. 

B. 

C. 

D. 

E. 

____ 1. This orbital notation shows an element with nine <u>total</u> electrons.

____ 2. This orbital notation shows an element with three <u>valence</u> electrons.

____ 5. This orbital notation shows an element with five <u>total</u> electrons.

____ 4. This orbital notation shows an element with the Lewis dot structure **shown below**.

Part IV: Multiple Choice

1. When the electron in an atom of hydrogen transitions from $n = 3$ to $n = 1$, which of the following are true?

   I. Energy is emitted.
   II. Energy is absorbed.
   III. The electron is now in its ground state.

   a. I only b. II only c. I and III only d. II and III only

2. What is the correct noble gas notation for the cation found within the compound $AlBr_3$?

   a. [Ne]$3s^2 3p^1$ c. [Ar]$4s^2 4p^6$
   b. [He]$2s^2 2p^6$ d. [Ar]$3s^2 3p^1$

3. Which of the following species has exactly two unpaired electrons in the ground state?

   a. Mg^{2+} b. Ti^{2+} c. Cr^{2+} d. Zn^{2+}

4. The ground state configuration of the Cr^{3+} ion is characterized by which of the following statements?

   I. Isoelectronic with a noble gas.
   II. An empty 4*s* orbital.
   III. Partially filled 3*d* orbitals.
   IV. The presence of unpaired electrons.

   a. II only b. II and III only c. II and IV only d. II, III and IV only

5. What is the correct electron configuration for the negatively charged anion found within the compound magnesium oxide?

   a. $1s^2 2s^2 2p^4$ c. $1s^2 2s^2 2p^6 3s^2$
   b. $1s^2 2s^2 2p^6$ d. $1s^2 2s^2 2p^6 3s^2 3p^4$

6. Which of the following, in their ground state, has exactly four unpaired electrons?

   I. Fe II. Cr III. O IV. Fe^{3+}

   a. I only b. III only c. II and IV only d. I and II only

7. Many of the unique properties of tin are due to the electron arrangement within the atom. What is the ground state electron configuration of tin?

   a. [Kr] $5s^2 5p^2$ c. [Kr] $5s^2 5d^{10} 5p^2$
   b. [Kr] $5s^2 4d^{10} 5p^2$ d. [Kr] $5s^2 4d^{10} 4f^{14} 5p^2$

8. Which of the following species has exactly three unpaired electrons in the ground state?

   a. Fe^{2+} b. P^{3-} c. Al^{3+} d. V^{2+}

9. All of the following pairs are isoelectronic EXCEPT:

   a. Cl^- and S^{2-} c. Ar and K^+
   b. Rb^+ and Xe d. Na^+ and N^{3-}

Unit 3 Part 5: Electromagnetic Radiation (EMR)

Electromagnetic radiation: (_____) is a form of _____ that will behave in some ways like a wave, but in some ways it will behave like stream of particles.

- **Photon:** a particle of electromagnetic radiation having zero mass and carrying a specific amount of energy.

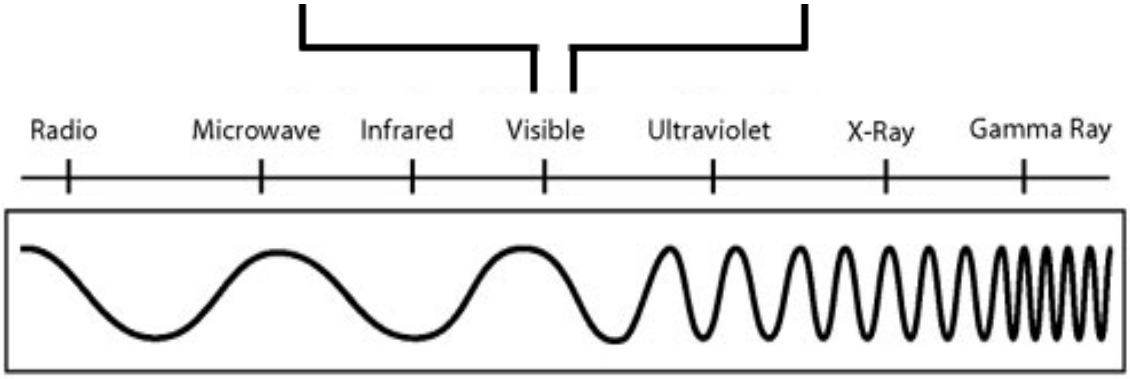
The relationship between energy, wavelength and frequency of any EMR wave can be expressed in 2 equations:

$$c = \lambda\nu \qquad E = h\nu$$

$$E = energy$$
$$\nu = frequency$$
$$\lambda = wavelength$$

Planck's constant, $h = 6.626 \times 10^{-34}$ J s

Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}



Guided Practice

1. A light sample is measure to have a wavelength of 8.27×10^{-7}m.
   a. What is the energy of the light in joules?
   b. What is the energy of the light in kJ/mol?

2. In an experiment, a molecule of hydrogen, H_2, absorbs a photon of electromagnetic radiation with a wavelength of 300 nm. The bond energy of H_2 gas is 432 kJ/mol.
    a. Calculate the frequency of the photons in Hz (sec^{-1}).

    b. Calculate the number of joules required to break the bond in a single molecule of H_2 gas.

    c. Does the photon have enough energy to break the bond in a molecule of H_2 gas? Mathematically justify your answer.

Independent Practice

3. A certain green light has a frequency of 3.13×10^{14} Hz.
    a. What is the energy of one photon of this light?

    b. What is the energy of this light measured in kilojoules/mole?

4. What is the wavelength, in nanometers, of a photon of light that has a frequency of 2.10×10^{14} Hz?

5. Which of the following are **directly** proportional? Which of the following are **inversely** proportional?
    a. energy and wavelength?
    b. wavelength and frequency?
    c. frequency and energy?

Multiple Choice Practice

6. The wavelength range for infrared radiation is 10^{-5} m, while that of ultraviolet radiation is 10^{-8} m. Which type of radiation has more energy, and why?
    a. Ultraviolet has more energy because it has a higher frequency.
    b. Ultraviolet has more energy because it has a longer wavelength.
    c. Infrared has more energy because it has a lower frequency.
    d. Infrared has more energy because it has a shorter wavelength.

7. Which of the following statements is or are true?
    1) An excited atom can return to a lower energy level by absorbing light energy.
    2) An atom can be excited by emitting light energy.
    3) As the energy of electromagnetic radiation increases, its frequency increases.
    4) The frequency and wavelength of light are inversely proportional.

    a. 1 and 2 b. 3 only c. 1 and 3 d. 3 and 4

8. What is the wavelength of yellow light having a frequency of 5.17×10^{14} s^{-1}?
    a. 3.60×10^{-10} m
    b. 5.80×10^{-7} m
    c. 1.55×10^{23} m
    d. 2.72×10^{-6} m

9. Which of the following statements is true?
    a. As energy increases, the frequency of the radiation decreases.
    b. As the wavelength of light increases, the frequency increases.
    c. Red light has a higher frequency than blue light.
    d. The product of wavelength and frequency of light in a vacuum is a constant.

Spectroscopy

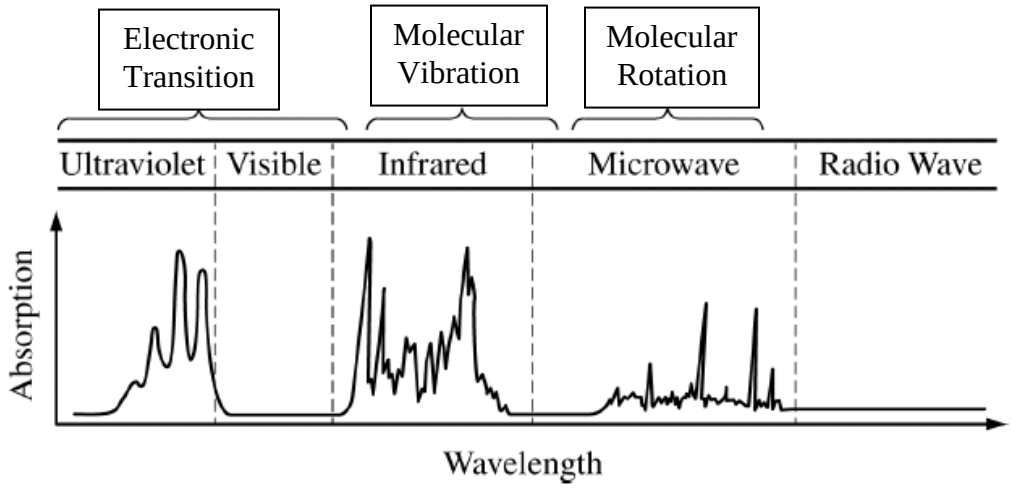
Spectroscopy: a study of the interaction between _____ and electromagnetic radiation

- Can be used to determine the atoms, molecules, or structure of a given substance
- So many kinds!!

Types of Spectroscopy

Name of Spectroscopy	Type of Radiation Used	Relative energy	What does it do to the atom/molecule?	What does it tell us?
Photoelectron spectroscopy (PES)	X-ray	very high	Removes electrons (valence and core).	• Identity of element • How tightly electrons are held by the nucleus
UV-Visible spectroscopy (UV-Vis)	Ultraviolet (UV)	high	Excites electrons to jump between energy levels.	• Identity of element or molecule
UV-Visible spectroscopy (Colorimetry)	Visible	medium		• Identity of element or molecule • Concentration of solution
IR (vibrational) spectroscopy	Infrared (IR)	low	Changes vibrations within bonds.	• Types of atoms, bonds, and functional groups within a molecule
Microwave (rotational) spectroscopy	Microwave	very low	Changes the rotation of atoms in bonds.	• Location of hydrogen atoms within a molecule

ABSORPTION SPECTRUM

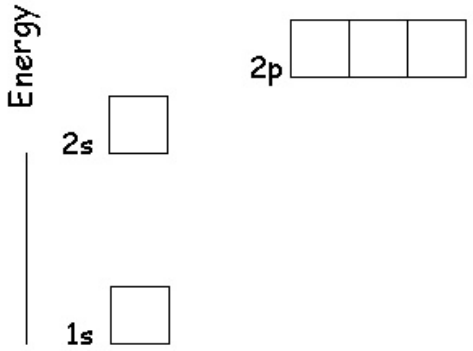
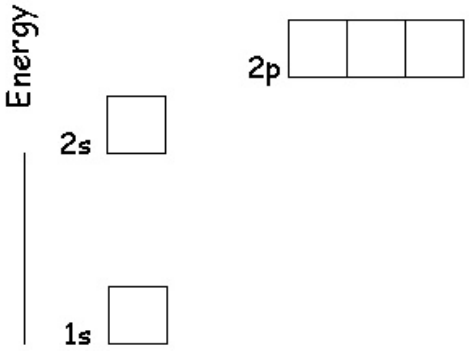


1. Based on the absorption spectrum shown above, rank the following transitions in order from least to greatest energy required: transition between vibrational states, between rotational states, and between electronic states.

2. Which type of spectroscopy have we performed in class? When, and for what purpose?

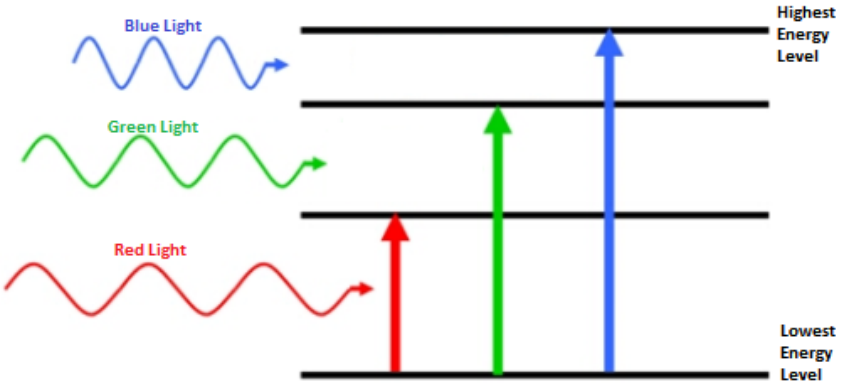
Atomic Absorption/Emission Spectra

Atoms exist in two states in relation to _____.

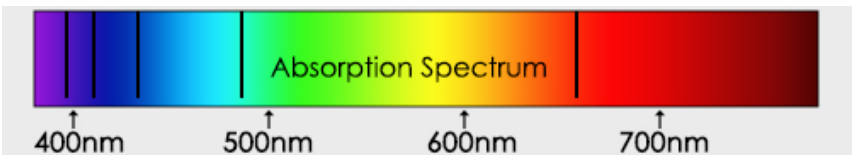
Ground State	Excited State – Added Energy!
Definition: electrons in their lowest energy state	Definition: **when energy is added,** an electron jumps up into a higher energy orbital
Example: lithium ground state $1s^2 2s^1$ 	Example: lithium excited state $1s^2 2p^1$ 

Atomic Absorption Spectra

- Used to determine the concentrations of _____ present in a given sample
- Electrons can move to a **higher** energy orbital by _____ a specific amount, or quantum, of energy.
- As an electron moves from its ground state to an excited state, the atom's potential energy _____.
- The _____ spectrum of an element is the relative intensity of each frequency of electromagnetic radiation absorbed by the atom as the atom's electrons jump **from the ground state to the excited state**.

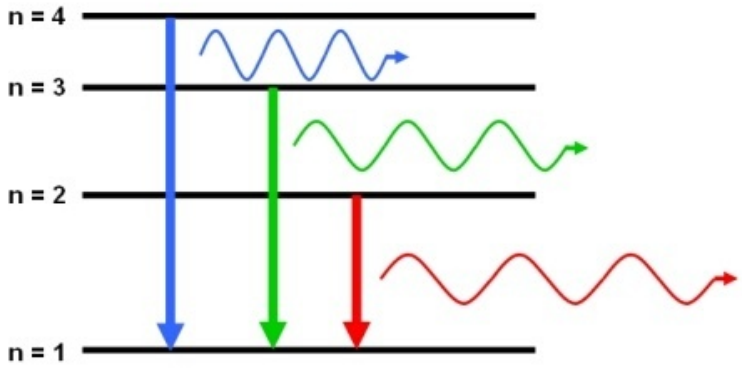


HYDROGEN SPECTRUM

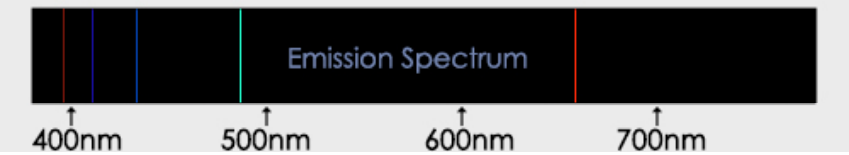


Atomic Emission Spectra (Bright-Line Spectra)

- Used to determine the identities of _____ present in a given sample
- When electrons fall from an **excited** state back to **ground** state, they _____ a specific amount, or quantum, of energy.
- The _____ spectrum of an element is the relative intensity of each frequency of electromagnetic radiation emitted (released) by the atom as the atom's electrons return from the excited state to the ground state.



HYDROGEN SPECTRUM



Let's Practice!

1. Using data from the emission spectra below, what is the most likely source of the unknown?

	Red	Orange	Yellow	Green	Blue	Violet
Lithium						
Sodium						
Potassium						
Calcium						
Strontium						
Barium						
Unknown						
λ (nm)	700nm	650nm	600nm	550nm	500nm	450nm

2. Which of the following hydrogen electron transitions will result in the absorption of light with the lowest energy?

   a. n = 2 to n = 3 b. n = 3 to n = 4 c. n = 3 to n = 2 d. n = 4 to n = 3

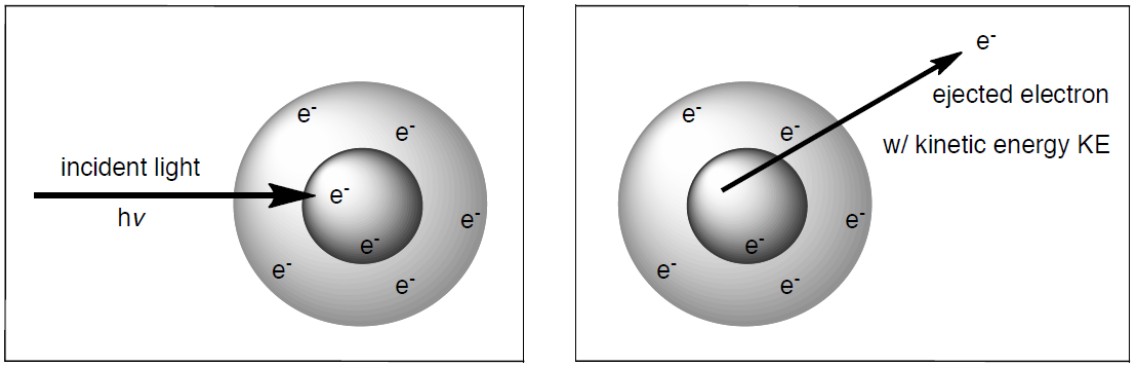
3. Which is an electron configuration of a fluorine atom in the excited state?

   a. $1s^2 2s^2 2p^4$ b. $1s^2 2s^2 2p^4 3s^1$ c. $1s^2 2s^2 2p^5$ d. $1s^2 2s^2 2p^5 3s^1$

Photoelectron Spectroscopy (PES)

Photoelectron Spectroscopy (_____): a technique to determine the ionization (or _____) energy of EVERY electron in an atom

- Ionization energy (for PES, more commonly referred to as the **binding energy**): the energy required to remove (_____) an electron from an atom
- For PES, the binding energy is commonly measured in kJ/mol or MJ/mol.
- Binding energy is plotted on the horizontal axis, with energy _____ (!!) from left to right (although sometimes this is flipped, so always check)

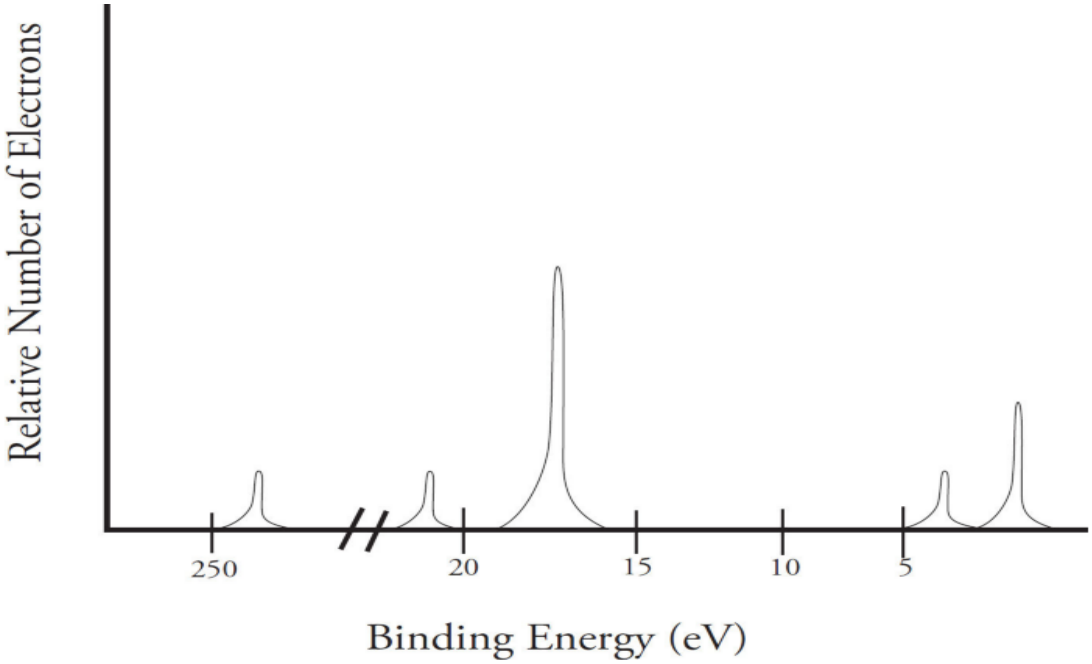


Incoming Radiant Energy = Binding Energy + kinetic energy (of ejected electron)

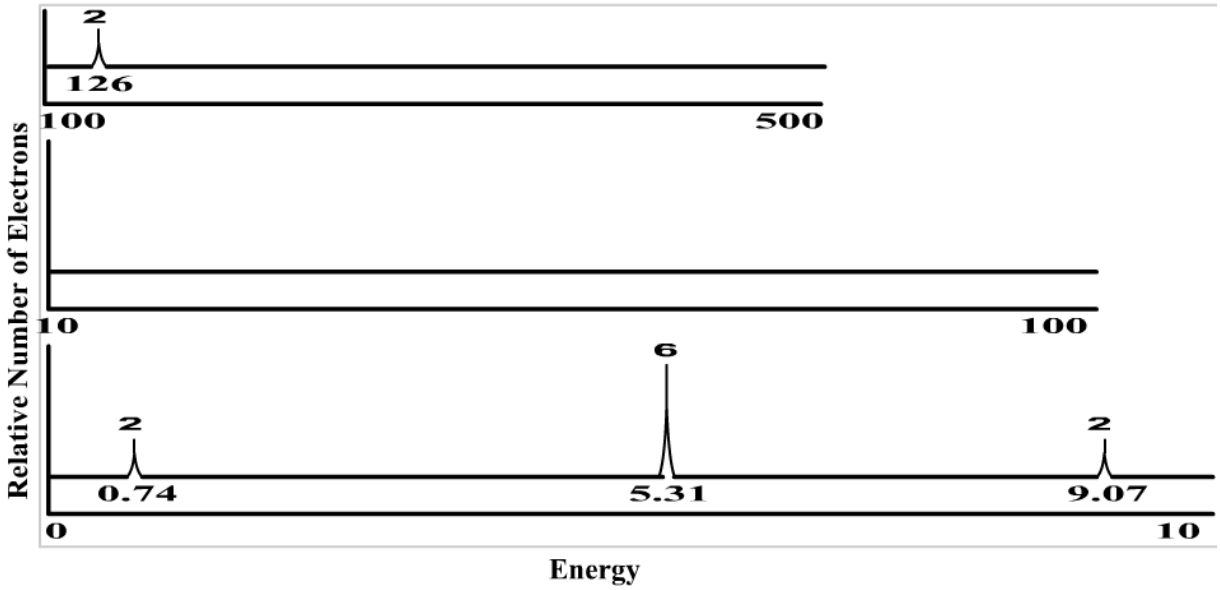
How to Interpret PES Spectra

- Peak _____ corresponds to the relative _____ of electrons in each sublevel of an atom
- Peak _____ corresponds to the relative amount of _____ required to remove each electron
  - Higher energies = sublevels found _____ to the nucleus (1s, 2s, etc)
  - Lower energies = sublevels found _____ from the nucleus
- When comparing PES from different atoms:
  - As the number of _____ in the nucleus increases, the binding energy will _____ for electrons in comprable sublevels
  - As the number of _____ in a specific sublevel increases, the peak height will _____ for electrons in comprable sublevels

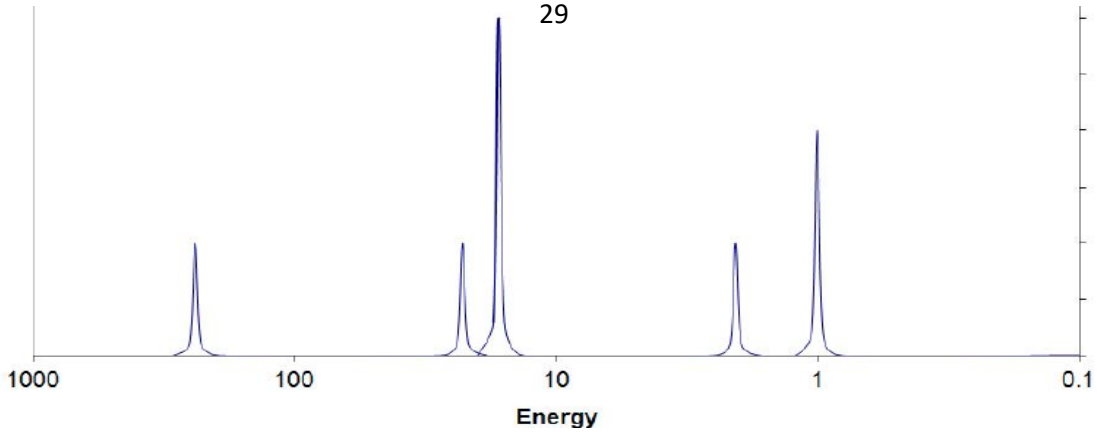
Guided Practice



1. The photoelectron spectrum of an element is shown above. Answer the following:
    a. Label the peaks with their corresponding orbital sublevel AND the number of electrons found.
    b. Identify the element represented by the photoelectron spectrum.
    c. Which electrons would be the first to be removed?
    d. Which electrons would be the last electrons to be removed?

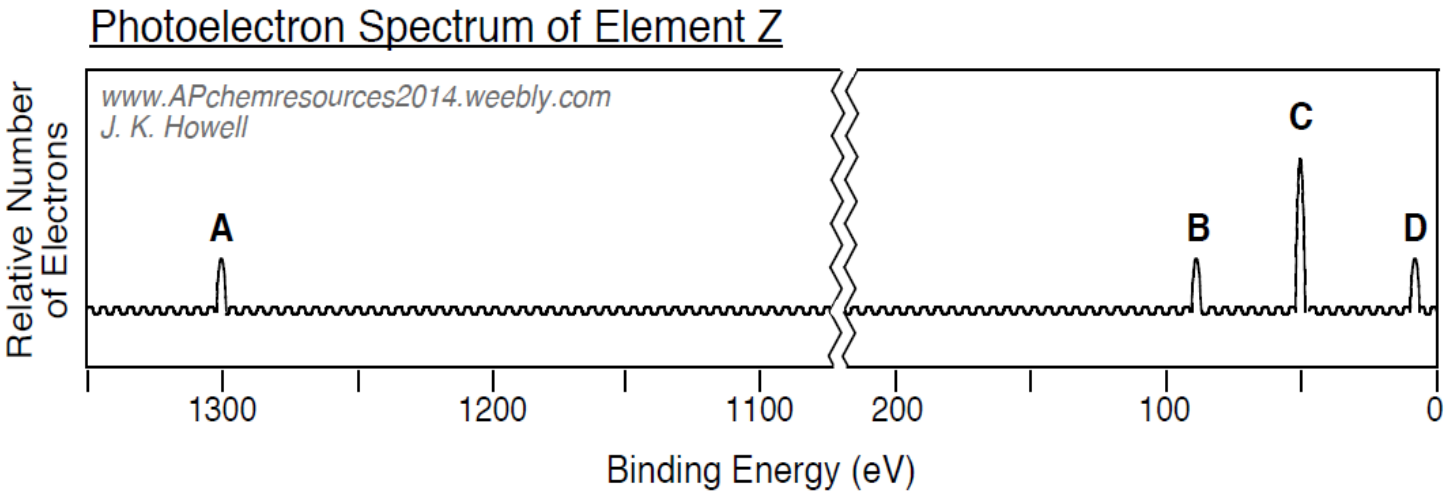


2. For the spectrum shown above,
    a. The element is _____.
    b. It has _____ valence electrons.
    c. Its noble gas configuration is:



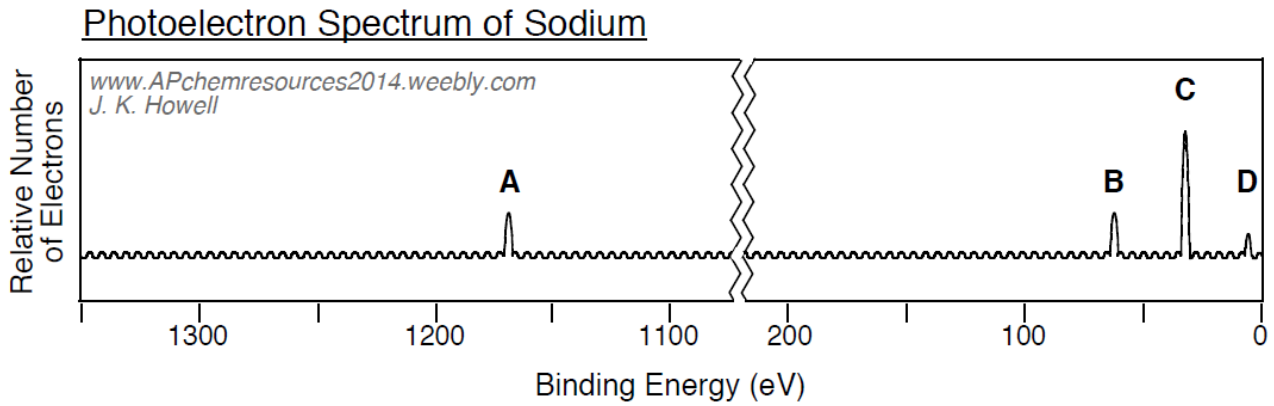
3. Above is the PES spectrum of sulfur.
    a. Label each peak in the spectrum to show which subshell it represents, and how many electrons are in it.
    b. On the spectrum, sketch in the relative locations and correct peak heights for the spectrum of oxygen, and label each with the corresponding orbital sublevel and the number of electrons found.
    c. Does it require more or less energy to remove a 2p electron from oxygen compared to a 2p electron from sulfur? Justify in terms of Coulomb's Law.

Independent Practice



4. What is the identity of element Z?
    a. Boron b. Carbon c. Neon d. Magnesium
5. Label the identity of each peak with principal quantum number, n (energy level), subshell (s, p, d, or f) and a superscript representing the number of electrons found in the subshell.
6. If the PES spectrum above had actually represented the element sodium, what would be different? List at least two differences you would expect to see:

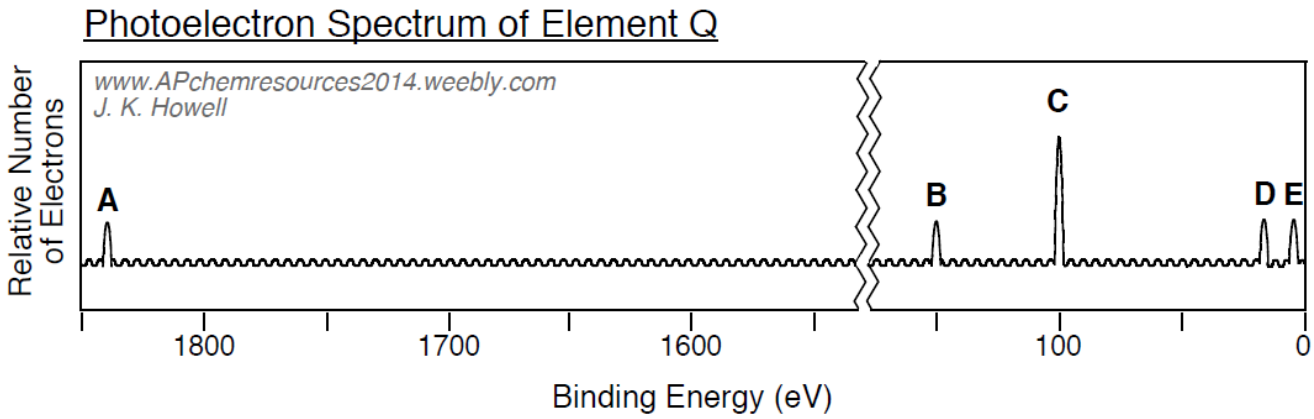
Questions 7 – 9 refer to the complete photoelectron spectrum of sodium shown below:



7. Label the identity of each peak with principal quantum number, n (energy level), subshell (s, p, d, or f) and a superscript representing the number of electrons found in the subshell.
8. Which peak (A-D) in the photoelectron spectrum of sodium represents an orbital containing 6 electrons? _____
9. Which peak (A-D) in the photoelectron spectrum of sodium shows electrons closest to the nucleus? _____

Multiple Choice Practice

Questions 10 – 13 refer to the complete photoelectron spectrum shown below:



10. The electrons that feel the strongest effective nuclear charge are given by which peak?
    a. A
    b. B
    c. E
    d. all electrons in this spectrum experience the same effective nuclear charge

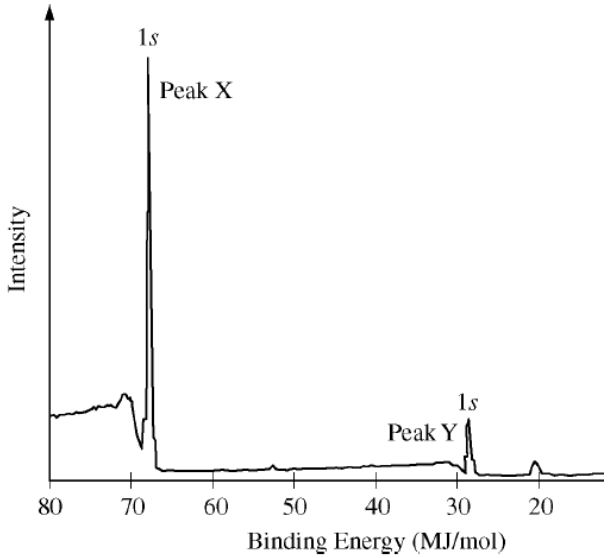
11. Which peaks in the photoelectron spectrum represent valence electrons of element Q?
    a. A only
    b. B and C
    c. D and E
    d. E only

12. Which peaks in the photoelectron spectrum of Q are given by the binding energy of *p* orbital electrons?
    a. C only
    b. D only
    c. C and E
    d. B, C, and D

13. What is the identity of element Q?
    a. carbon
    b. oxygen
    c. magnesium
    d. silicon

14. When a given sample was analyzed, it was determined to contain only atoms of boron and oxygen. The portion of the PES spectrum showing only the 1s peaks is provided. Which peak belongs to boron and why?

    a. Peak X, because B is less electronegative than O.
    b. Peak X, because B has a larger atomic radius than O.
    c. Peak Y, because B has a smaller nuclear charge than O.
    d. Peak Y, because B has a greater first ionization energy than O.

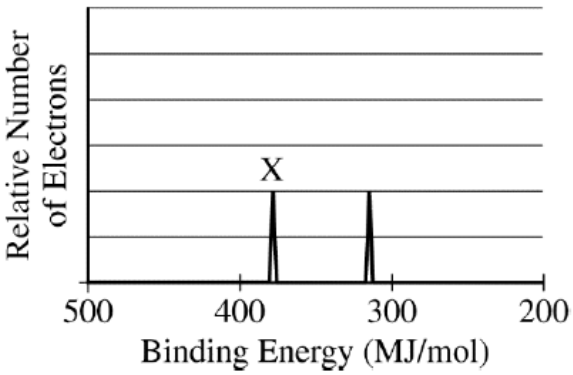


15. Atoms of four elements are examined: carbon, nitrogen, neon and sulfur. Which element would have a photoelectron spectrum in which the peak representing electrons with the lowest ionization energy would be three times higher than all other peaks?

    a. Carbon b. Nitrogen c. Neon d. Sulfur

16. The PES spectra of the 1s electrons for two isoelectronic species, Sr^{2+} and Kr, is shown to the right. Which species resulted in peak X, and why?

    a. Kr, because it has a completely filled valence shell.
    b. Kr, because it has a smaller radius than Sr^{2+}.
    c. Sr^{2+}, because it has a greater mass than Kr.
    d. Sr^{2+}, because it has more protons than Kr.



17. An atom of oxygen is in an excited state. When an electron in this atom moves from the third to the second energy level, energy is:

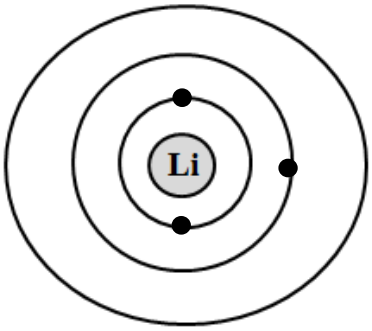
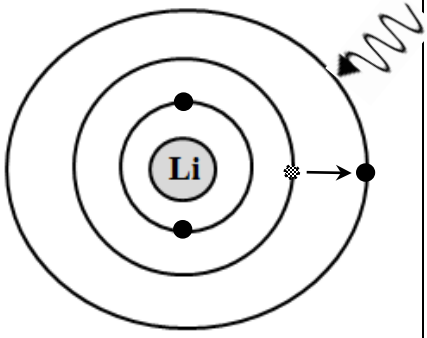
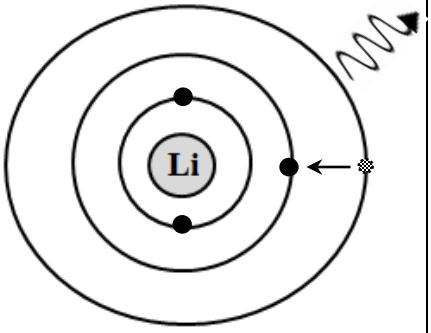
    a. emitted by the nucleus c. absorbed by the nucleus
    b. emitted by the electron c. absorbed by the electron

Flame Test Lab

Introduction

The energy levels in atoms are the key to the production and detection of light. Energy levels exist for electrons in atoms. The colors result from electrons falling back from excited levels to lower energy levels. Observations of light emitted by the elements are also evidence for the existence of energy levels, and the sub-levels within the shells. Different elements emit different emission spectra when they are excited because each type of element has a unique energy level system. Each element has a different set of emission colors because each element has a different arrangement of electrons, and a different number of protons in the nucleus to attract those electrons. We will then identify an unknown element by comparing the color of the unknown with the flame color of our known elements.

Light Bright: Making Colors with Chemistry

 **Step 1**: Ground State	 **Step 2**: Add Energy, Get Excited!	 **Step 3**: Return to Ground State
• Electrons found at lowest energy level.	• Energy added: o fire (heat) o UV light (black lights) • Electron jumps up to higher energy level	• Electron falls down to lower energy level. • Energy released, sometimes in the form of light! (pretty…)

Chemicals

1.0 M lithium chloride, LiCl, solution
1.0 M potassium chloride, KCl, solution
1.0 M calcium chloride, $CaCl_2$, solution
1.0 M strontium chloride, $SrCl_2$, solution
1.0 M copper chloride, $CuCl_2$, solution
1.0 M barium chloride, $BaCl_2$, solution

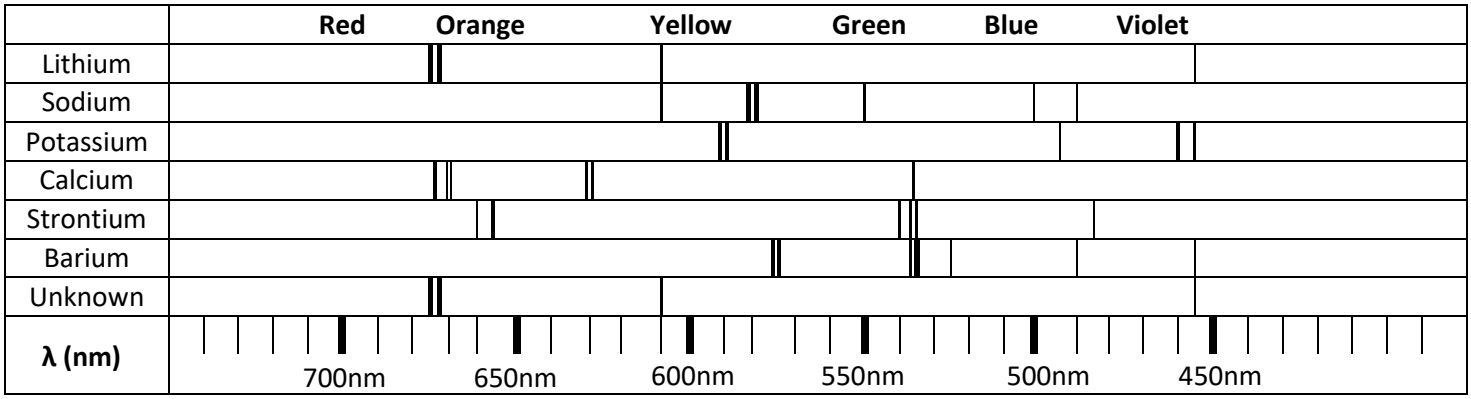
Procedure

1. Using a beaker, obtain one of the wooden splints that have been soaking in the beakers of metal salts at the front of the room. Use your beaker to carry the wooden splint from the front of the room to your lab station. This will avoid dripping chemical around the room.
2. Light your Bunsen burner. It is IMPORTANT that one member of your lab group watches the burner at ALL TIMES.
3. Insert the wooden splint ***into the flame*** of the Bunsen burner.
4. Observe the color of the flame produced with your naked eye and record this color on your lab worksheet.
5. Place each used wooden splint into a beaker of tap water after removing it from the flame to ensure it is extinguished.
6. Repeat steps 1-6 with each of the different metal salts.
7. When you have finished using the Bunsen burner, ensure that the gas valve is completely in the OFF position.

Data Table 1: Characteristic Colors of Flame

Ion	Color of Flame	Noble Gas Electron Configuration
Lithium		
Barium		
Potassium		
Calcium		
Strontium		
Copper		
Unknown		

Data Table 2: Emission Spectra of Elements



Analysis:

1. What is the identity of your unknown metal cation? How can you tell? (Hint: there are two reasons, and only one of those reasons is based on your lab data!)

Error Analysis:

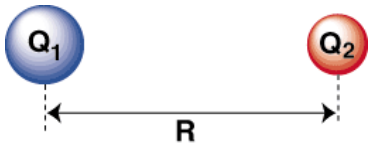
What possible sources of error exist in this lab that might lead to an incorrect identification of your unknown? Be specific!

Periodic Trends

Periodic Trends are specific _____ that are present in the periodic table that illustrate different aspects of a certain element.

Almost all of the properties that are asked about in exam questions rely on Coulombic (electrostatic) attraction between outer electrons and the nucleus. Hence, you should refer to Coulomb's law in your answers!

$$E \propto \frac{Q_1 Q_2}{r}$$



E = ionization energy, the energy needed to _____ the outermost electron.
Q_1 = charge of an electron, -1.
Q_2 = effective nuclear charge (_____) of protons in nucleus.
r = distance between charged particles, which can be approximated by the period (n, energy level).

In short, the **energy of attraction or repulsion** between charged particles is:

- _____ proportional to the **magnitude** (size) of the charges
- _____ proportional to the **distance** between the charges

All Periodic Trends can be understood in terms of Coulomb's Law(!)

1. Electrons are attracted to the protons in the nucleus of an atom.
    a. The _____ an electron is to the nucleus, the **MORE** strongly it is attracted.
        i. Distance between electrons and the nucleus can be approximated by ____ (main energy level)
    b. The _____ protons in a nucleus, the **MORE** strongly an electron is attracted.
2. Electrons are repelled by other electrons in an atom.
    a. If other electrons are _____ a valence electron and the nucleus, the valence electron will be **LESS** strongly attracted to the nucleus (this is known as _____).

Effective Nuclear Charge (_____)**:** net positive charge experienced by electrons

→ Attraction of a valence electron to a proton is partially shielded by the inner shell (_____) electrons, and so the farther away an electron is from the nucleus, the less _____ charge it feels from the nucleus!

→ Thus, following Coulomb's Law:

- The nucleus of atoms with a **higher Z_{eff}** (at the same energy level, n) will be _____ attractive to their valence electrons
- The nucleus of atoms with a **lower Z_{eff}** (at the same energy level, n) will be _____ attractive to their valence electrons

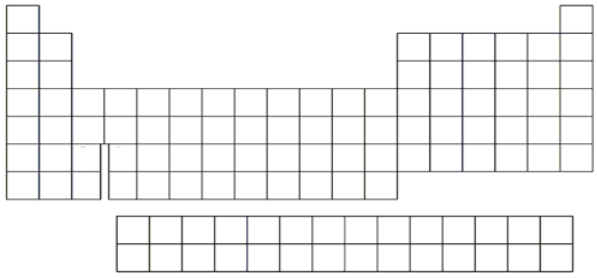
→ At the _____ energy level, Z_{eff} is directly proportional to the number of _____!

NOTE: Energy level trumps Z_{eff}!
If you're comparing elements that have different n AND different Z_{eff}, ____ matters more!!

Trendy Trends: Periodic Trends to Know

Atomic radius (size of atom): distance between the _____ and valence electrons.

Hint: Snowman!



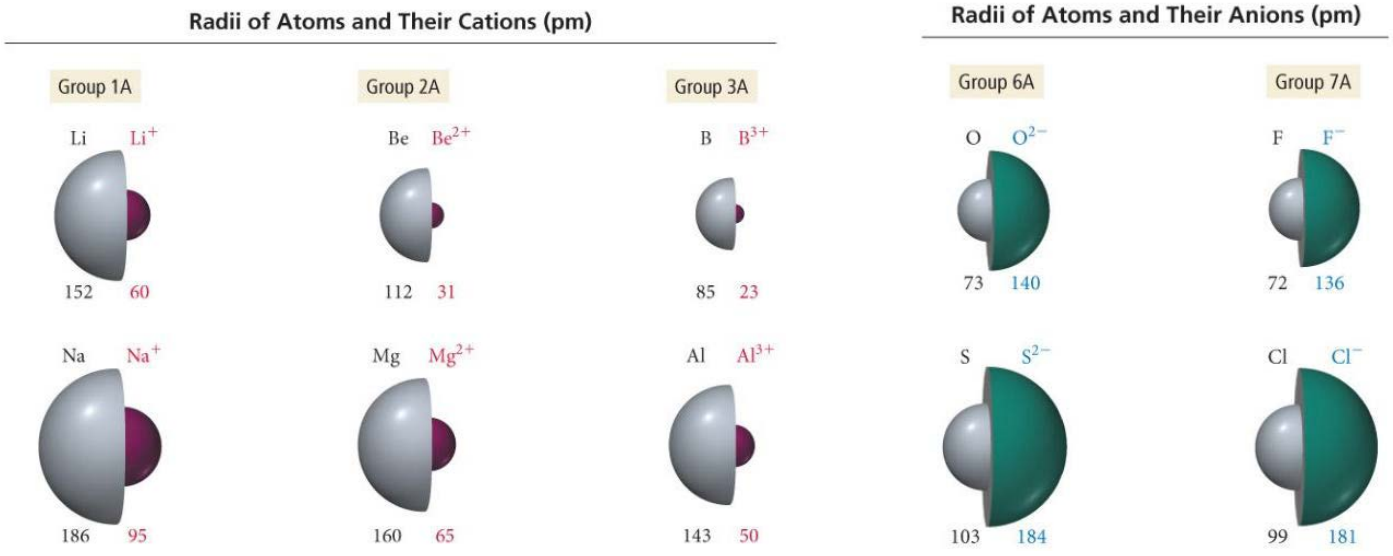
Ionic radius: distance from the nucleus to valence electrons in a charged _____.

Most comparisons between an atom and their ion or ions of the same atom can be explained by **e^-/e^- repulsion**:

- Positively charged cations are **SMALLER** than the neutral atom because _____ electrons in the outermost shell results in _____ e^-/e^- repulsion, thus valence electrons are _____ to the nucleus.
- Negatively charged anions are **LARGER** than the neutral atom because _____ electrons in the outermost shell results in _____ e^-/e^- repulsion, thus valence electrons are _____ from the nucleus.

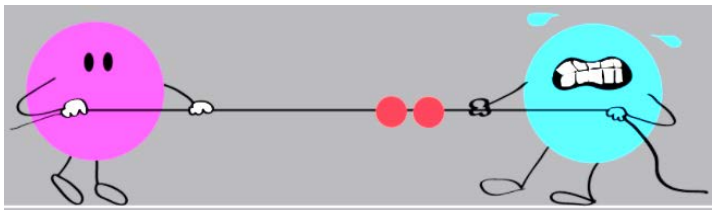
Only one type of comparison between an atom and their ion(s) should **NOT** utilize the e^-/e^- repulsion argument:

- Metals cations which have lost sufficient electrons such that their valence electrons are now in a _____ energy level (n). Examples: Sr vs Sr^{2+}, Na vs Na^+, Al^{2+} vs Al^{3+}, etc.
- Why? Because **n matters more than Z_{eff}!!!** If a species has their outermost electrons on a lower energy level (n), their valence electrons are _____ to and thus _____ attracted to the nucleus.



Electronegativity: attraction of an atom for pair of _____ level electrons in a covalent bond with another atom.

Think of the atoms as playing "tug of war" with their _____ shell electrons!



Electronegativity of an atom is determined by combining the following two trends:

Ionization Energy vs Electron Affinity

Ionization Energy	Electron Affinity
Energy required to remove an electron	Energy change when electron is added to an atom
$X(g) + \text{energy} \rightarrow X^+(g) + e^-$	$X(g) + e^- \rightarrow X^-(g) + \text{energy}$
endothermic (+kJ/mol)	first electron added is always exothermic (−kJ/mol)

Higher attraction between nucleus and electron = _____ ionization energy and _____ electron affinity!

Metallic vs Non-Metallic Character

Metallic Character	Non-Metallic Character
Metals react by losing electrons	Non-metals react by gaining electrons
How easy it is to remove an electron from an atom or ion	How hard is it to remove an electron from an atom or ion
___ Ionization Energy = ___ Metallic Character	___ Ionization Energy = ___ Non-Metallic Character

Let's Practice!

O Oxygen	F Fluorine
S Sulfur	Cl Chlorine

Consider the elements shown above. Which element has the:

a) Highest ionization energy?_____

b) Highest electron affinity?_____

c) Highest Metallic Character?_____

d) Highest Non-Metallic Character?_____

Successive Ionization Energies

Removing one electron, and then another electron, and then another electron…

- 1st Ionization Energy: (_____) energy required to remove the first (highest energy) electron
- 2nd Ionization Energy: (_____) energy required to remove the second electron (second highest energy)
- Each additional electron requires **MORE** energy to remove than the previous one:

$$IE_1 < IE_2 < IE_3 \text{ (etc.)}$$

- Valence (outer) electrons require much **LESS** energy to remove than core (inner) electrons

$$IE_{valence} << IE_{core}$$

TABLE 8.1 Successive Values of Ionization Energies for the Elements Sodium through Argon (kJ/mol)

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	496	4560					
Mg	738	1450	7730		Core electrons		
Al	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,100		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
Cl	1251	2300	3820	5160	6540	9460	11,000
Ar	1521	2670	3930	5770	7240	8780	12,000

You can identify an element from the pattern of successive ionization energies!

FR Practice

IE_1	IE_2	IE_3	IE_4	IE_5
801 kJ/mol	2,426 kJ/mol	3,660 kJ/mol	24,682 kJ/mol	32,508 kJ/mol

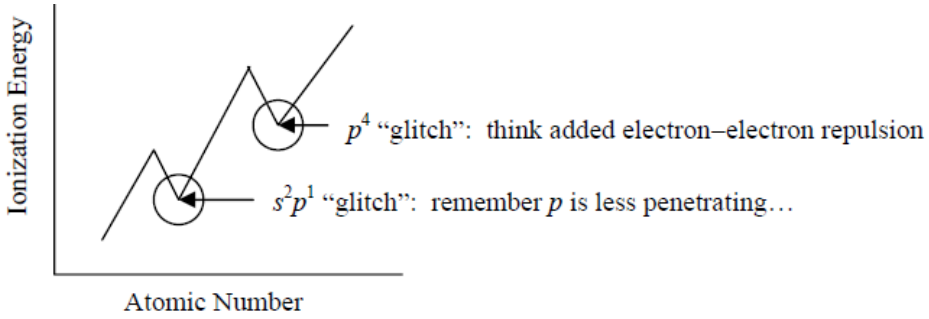
1. Which element from Period 2 does the table of successive ionizations energies above represent? Explain.

2. Would it require more energy to remove an electron from an F^+ ion or an F^{2+} ion? Justify your answer using Coulomb's Law.

Ionization Energy Exceptions:

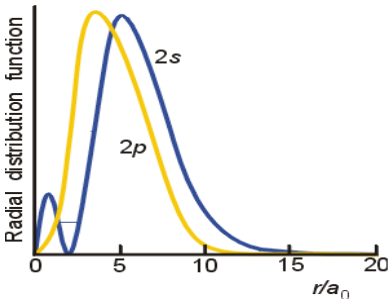
You MUST know these!

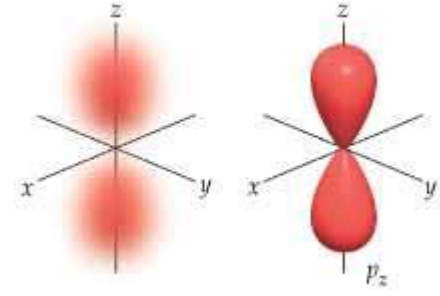
*these exceptions only matter when both elements are found on the same period (energy level).

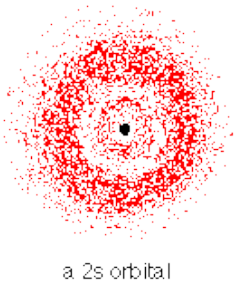


1. **Group 5 vs Group 6: "p^4 glitch"** (Example: oxygen vs nitrogen): It's easier to remove an electron from oxygen ($2p^4$) because of the _____ of the paired electrons (whereas the $2p^3$ electrons in nitrogen (Group 5) are all unpaired).

2. **Group 2 vs Group 3: "s^2p^1 glitch"** (Example: beryllium vs boron): It's easier to remove an electron from boron ($2s^22p^1$) because it's being removed from the p orbital, and p orbitals do not _____ the nuclear region as greatly as the s orbitals, so electrons in a p orbital are not as tightly held.





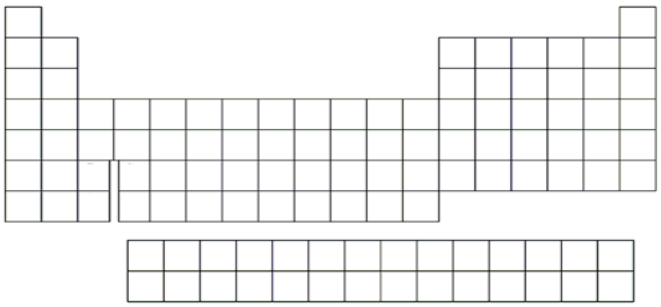


s orbital: higher nuclear penetration more electron density in greater proximity to nucleus

BE CAREFUL! This is NOT about distance between electron and nucleus - on average, electrons in a 2s orbital and 2p orbital are the same distance from the nucleus.

Ionization Energy Trend

Hint: Ice cream!

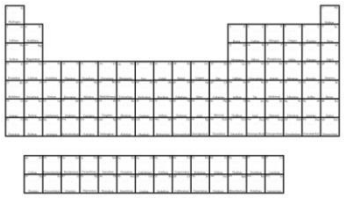
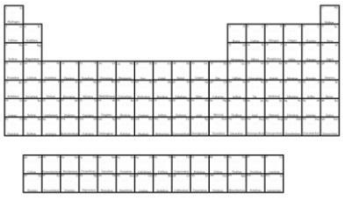
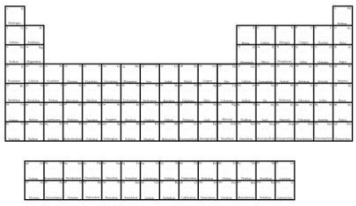
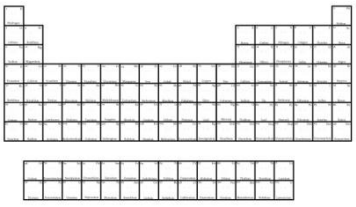
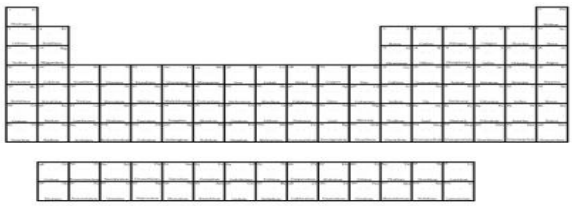


One last trend:

- **<u>Reactivity</u>** depends on whether the element reacts by losing electrons (metals) or gaining electrons (nonmetals).
    - <u>Metals are MORE reactive as you move down a column</u>: because metals _____ electrons as they react, **LESS** attraction between valence electrons and the nucleus will result in a _____ reactive metal.
    - <u>Non-metals are LESS reactive as you move down a column</u>: because non-metals _____ electrons as they react, **LESS** attraction between valence electrons and the nucleus will result in a _____ reactive non-metal.

In Summary

Draw arrows in the direction that each trend increases!

Atomic Radius Trend 	Metallic Character Trend 
Electronegativity Trend 	Ionization Energy Trend 
Reactivity Trend 	

Multiple Choice Practice

	Ionization Energy
First	740 kJ/mol
Second	1,450 kJ/mol
Third	7,730 kJ/mol
Fourth	10,540 kJ/mol
Fifth	13,610 kJ/mol

1. Given the table of ionization energies for the unknown element M shown above, which of the following is the most probable empirical formula for the compound element M would form with fluorine?

   a. MF b. MF_2 c. MF_3 d. MF_4

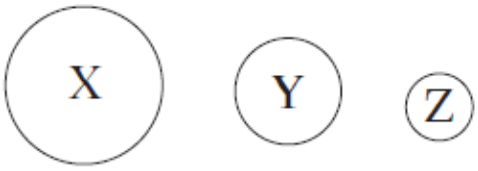
2. Why does an ion of phosphorus, P^{3-}, have a larger radius than a neutral atom of phosphorus?

   a. There is a greater Coulombic attraction between the nucleus and the electrons in P^{3-}.
   b. The core electrons in P^{3-} exert a weaker shielding force than those of a neutral atom.
   c. The nuclear charge is weaker in P^{3-} than it is in P.
   d. The electrons in P^{3-} have a greater Coulombic repulsion than those in the neutral atom.

3. Which of the following, sodium or magnesium, has greater metallic character and why?

   a. Sodium, because sodium has fewer principal energy levels than magnesium.
   b. Sodium, because sodium has a lower first ionization energy than magnesium.
   c. Magnesium, because of the repulsion of magnesium's paired 4s electrons.
   d. Magnesium, because magnesium has a greater effective nuclear charge than sodium.

4. The diagram below shows the relative atomic sizes of three different elements from the same period. Which of the following statements must be true if trend exceptions are ignored?



   a. The effective nuclear charge will be the greatest in element X.
   b. The first ionization energy will be greatest in element X.
   c. The electron shielding effect will be greatest in element Z.
   d. The electronegativity value will be greatest in element Z.

5. Which of the following isoelectric species has the smallest radius?

   a. S^{2-} b. Cl^{-} c. Ar d. K^{+}

6. Nitrogen's electronegativity value is between those of phosphorus and oxygen. Which of the following correctly describes the relationship between the three values?

    a. The value for nitrogen is less than that of phosphorus because nitrogen is larger, but greater than that of oxygen because nitrogen has a greater effective nuclear charge.
    b. The value for nitrogen is less than that of phosphorus because nitrogen has fewer protons, but greater than that of oxygen because nitrogen has fewer valence electrons.
    c. The value for nitrogen is greater than that of phosphorus because nitrogen has fewer electrons, but less than that of oxygen because nitrogen is smaller.
    d. The value for nitrogen is greater than that of phosphorus because nitrogen is smaller, but less than that of oxygen because nitrogen has a smaller effective nuclear charge.

7. Which of the following, argon or krypton, has a higher IE_1 and why?

    a. Argon, because argon has fewer principal energy levels than krypton.
    b. Argon, because argon has a larger effective nuclear charge than krypton.
    c. Krypton, because krypton has more principal energy levels than argon.
    d. Krypton, because krypton has a larger effective nuclear charge than argon.

8. The first five ionization energies for an element are listed in the table below.

First	Second	Third	Fourth	Fifth
8 eV	15 eV	80 eV	109 eV	141 eV

    Based on the ionization energy table, the element is most likely to be

    a. Sodium b. Magnesium c. Aluminum d. Silicon

9. Which of the following statements is true regarding sodium and chlorine?

    a. Sodium has a greater electronegativity and a larger first ionization energy.
    b. Sodium has a larger first ionization energy and a larger atomic radius.
    c. Chlorine has a larger atomic radius and a greater electronegativity.
    d. Chlorine has a greater electronegativity and a larger first ionization energy.

10. Consider the halogens chlorine and bromine. Which has a larger atomic radius and why?

    a. Chlorine has a larger atomic radius because it has an increased number of principal energy levels.
    b. Chlorine has a larger atomic radius because it has a higher effective nuclear charge.
    c. Bromine has a larger atomic radius because it has an increased number of principal energy levels.
    d. Bromine has a larger atomic radius because it has a higher effective nuclear charge.

11. Which of the following species is NOT isoelectronic to Br^-?

    a. Se^{2-} b. Kr c. Rb^+ d. K^+

How to Answer Periodic Trends Free Response Questions

Justifying all of the trends on the periodic table can be simplified using these two generalizations:

1. Use ***number of protons*** *(or Z_{eff})* to justify trends **across a period**.
2. Use **increased distance (greater value of *n*)** to justify trends **down a group**.

How to Earn Full Points on Periodic Trends Problems
Follow these three steps EVERY time you answer a periodicity question!

1) Locate *both* elements on the periodic table and **state** the principal energy level (*n*) and the sublevel containing the valence electrons for *each* element.
2) Do they have the same or different ___ values?
3) If same *n*: argue with number of _____; if different *n*: argue with *n* vs. *n* _____.

REMEMBER: a trend is not an explanation!

Simply identifying a trend (atomic radius decreases as you move from left to right across a period, electronegativity decreases as you move down a column, etc) earns _____ points!

Avoid Losing Easy Points

1. When explaining, you must refer to **ALL** species (atoms, ions) referenced in the question, or you will not get full credit.
2. Read the question: justify with "principles of atomic structure" or "Coulomb's Law" (it will always be one or the other ☺).

Specific Question Types

1. Comparisons between _____ species: explain with **number of p^+**
   a. Isoelectronic species with _____ protons are **SMALLER** because the valence electrons are _____ attracted to and thus **CLOSER** to the nucleus.
   b. Isoelectronic species with _____ protons are **LARGER** because the valence electrons are _____ attracted to and thus **FARTHER** from the nucleus.
2. Comparisons between an atom and its ion/ions of the same atom, _____ n: explain with **e^-/e^- repulsion**
   a. Positively charged cations are **SMALLER** than the neutral atom because of_____ e^-/e^- repulsion, thus valence electrons are **CLOSER** to the nucleus.
   b. Negatively charged anions are **LARGER** than the neutral atom because of_____ e^-/e^- repulsion, thus valence electrons are **FARTHER** from the nucleus.
3. Comparisons between an atom and its ion/ions of the same atom, _____ n: explain with **distance**
   a. If a species has their outermost electrons on a lower energy level (n), their valence electrons are _____ to and thus _____ attracted to the nucleus.

Guided Practice: How will an answer be different when explained in terms of "atomic structure" vs "Coulomb's Law"?

1. Which ion would require more energy to remove an electron: Mn^{3+} or Mn^{4+}? Explain using principles of atomic structure.

2. Which ion would require more energy to remove an electron: Mn^{3+} or Mn^{4+}? Be sure to reference Coulomb's Law in your explanation.

Independent Practice

1. Using principles of atomic structure, explain why the Na^{+} ion is larger than the Li^{+} ion.

2. Which has a higher metallic character, an atom of lithium or an atom of boron? Explain your answer in terms of Coulomb's Law.

3. Which would be greater for the aluminum atom, its first ionization energy or its second ionization energy? Explain using principles of atomic structure.

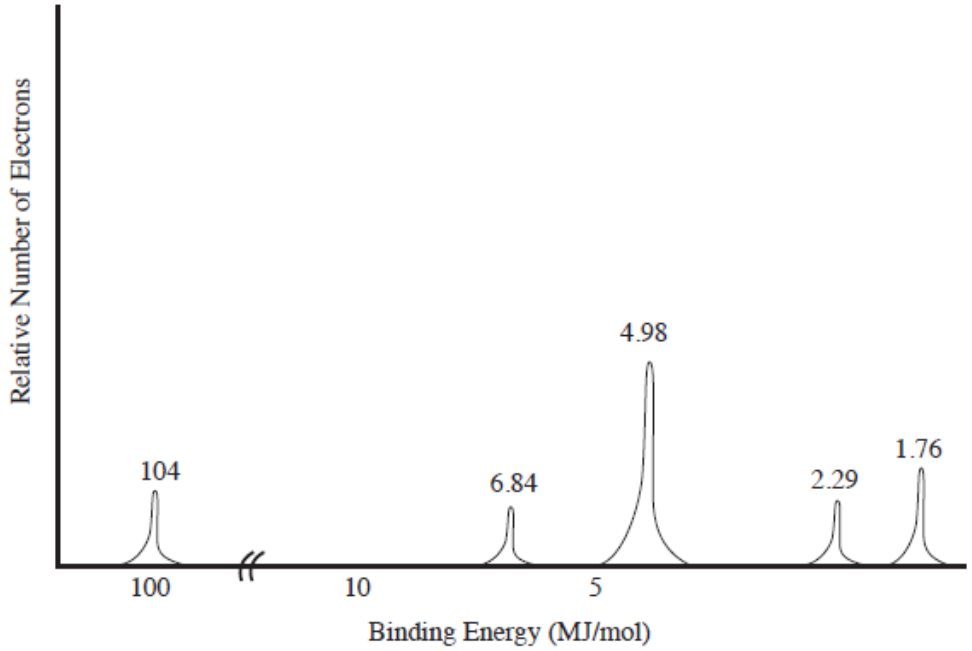
4. Which element, Cl or Br, has a greater first ionization energy? Justify your answer using Coulomb's Law.

Unit 10 Multiple Choice Practice

1. If europium (Eu) consists of two naturally occurring isotopes, ^{151}Eu with a percent abundance of 48.05% and ^{153}Eu with a percent abundance of 51.97%, what is the average atomic mass of europium?

   a. 152.00 amu
   b. 152.48 amu
   c. 151.96 amu
   d. 152.04 amu

Use the PES spectra below to answer questions 2-5.



2. What element does this spectrum represent?

   a. Boron
   b. Nitrogen
   c. Aluminum
   d. Phosphorus

3. Which peak represents the 2s subshell?

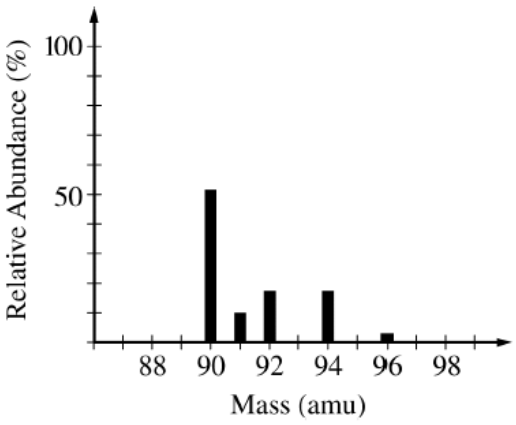
   a. The peak at 104 MJ/mol
   b. The peak at 6.84 MJ/mol
   c. The peak at 2.29 MJ/mol
   d. The peak at 1.76 MJ/mol

4. An electron from which peak would have the greatest velocity after ejection?

   a. The peak at 104 MJ/mol
   b. The peak at 6.84 MJ/mol
   c. The peak at 4.98 MJ/mol
   d. The peak at 1.76 MJ/mol

5. How many valence electrons does this atom have?

   a. 2
   b. 3
   c. 4
   d. 5



6. Given the mass spectrum on the left of an unknown element, which of the following is a true statement about this element?
   a. The element can be identified as a transition metal, and each peak represents a possible oxidation state.
   b. The element must have five distinct electron subshells (or sublevels).
   c. The element's atomic mass is 90 amu.
   d. The element's atomic mass falls between 90-92 amu.

7. The effective nuclear charge experienced by the outermost electron of Na is different than the effective nuclear charge experienced by the outermost electron of Ne. This difference best accounts for which of the following?
   a. Na has a greater density at standard conditions than Ne.
   b. Na has a lower first ionization energy than Ne.
   c. Na has a higher neutron-to-proton ratio than Ne.
   d. Na has fewer naturally occurring isotopes than Ne.

8. What of the following represents the ground state electron configuration for the Mn^{3+} ion?
   a. $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^4$
   b. $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^4\ 4s^2$
   c. $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^2\ 4s^2$
   d. $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^3\ 4s^1$

9. Which neutral atom of the following elements would have the most unpaired electrons?
   a. Titanium b. Manganese c. Nickel d. Zinc

10. What is the energy of a photon of yellow light having a wavelength of 5.80×10^{-7} m?
    a. 3.42×10^{-19} J
    b. 3.84×10^{-27} J
    c. 1.14×10^{-27} J
    d. 1.74×10^{3} J

11. The first ionization energy for a neutral atom of chlorine is 1.25 MJ/mol and the first ionization energy for a neutral atom of argon is 1.52 MJ/mol. How would the first ionization energy for a neutral atom of potassium compare to those values?

    a. It would be greater than both because potassium carries a greater nuclear charge than either chlorine or argon.
    b. It would be greater than both because the size of a potassium atom is smaller than an atom of either chlorine or argon.
    c. It would be less than both because there are more electrons in potassium, meaning they repel each other more effectively and less energy is needed to remove one.
    d. It would be less than both because a valence electron of potassium is farther from the nucleus than one of either chlorine or argon.

12. What is the most likely electron configuration for a sodium ion?

    a. $1s^2 2s^2 2p^5$
    b. $1s^2 2s^2 2p^6$
    c. $1s^2 2s^2 2p^6 3s^1$
    d. $1s^2 2s^2 2p^6 3s^2$

13. An atom of silicon in its ground state is subjected to a frequency of light that is high enough to cause electron ejection. An electron from which subshell of silicon would have the highest kinetic energy after ejection?

    a. 1s b. 2p c. 3p d. 4s

14. Examining data obtained from mass spectrometry supports which of the following?

    a. The common oxidation states of elements
    b. Atomic size trends within the periodic table
    c. Ionization energy trends within the periodic table
    d. The existence of isotopes.

15. In general, do metals or nonmetals from the same period have higher ionization energies? Why?

    a. Metals have higher ionization energies because they usually have more protons than nonmetals.
    b. Nonmetals have higher ionization energies because they are larger than metals and harder to ionize.
    c. Metals have higher ionization energies because their valence electrons are further from the nucleus than those of nonmetals.
    d. Nonmetals have higher ionization energies because their valence electrons experience greater Coulombic attraction to the nucleus than those of metals.

16. Which of the following ions would have the most unpaired electrons?

    a. Mn^{2+} b. Ni^{3+} c. Ti^{2+} d. Cr^{6+}

Use the following information to answer questions 17–19.

The radius of atoms and ions is typically measured in Angstroms (Å), which is equivalent to 1×10^{-10} m. Below is a table of information for three different elements.

Element	Atomic Radius (Å)	Ionic Radius (Å)
Ne	0.38	n/a
P	0.98	1.00
Zn	1.42	1.35

17. The phosphorus ion is larger than a neutral phosphorus atom, yet a zinc ion is smaller than a neutral zinc atom. Which of the following statements best explains why?
    a. The zinc atom has more protons than the phosphorus atom.
    b. The valence electrons in a phosphorus atom are closer to the nucleus than those of a zinc atom.
    c. Phosphorus gains electrons when forming an ion, but zinc loses them.
    d. Phosphorus has a greater electronegativity than zinc.

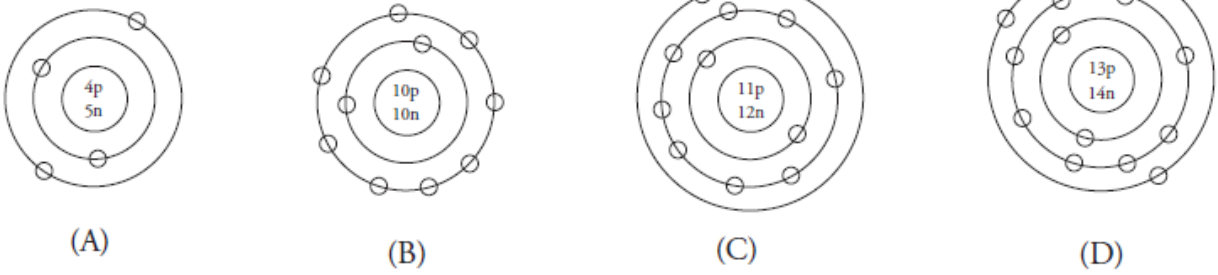
18. Neon has a smaller atomic radius than phosphorus because:
    a. Unlike neon, phosphorus has electrons present in its third energy level.
    b. Phosphorus has more protons than neon, which increases the repulsive forces in the atom.
    c. The electrons in a neon atom are all found in a single energy level.
    d. Phosphorus can form anions, while neon is unable to form any ions.

19. Which of the following represents the correct electron configuration for the zinc ion, Zn^{2+}?
    a. $[Ar]3d^{10}$
    b. $[Ar]4s^2 3d^8$
    c. $[Ar]4s^2 4d^8$
    d. $[Ar]4d^{10}$

20. Which of the following, calcium or strontium, has greater metallic character and why?
    a. Calcium, because calcium has fewer principal energy levels than strontium.
    b. Calcium, because calcium has a lower effective nuclear charge than strontium.
    c. Strontium, because strontium has a greater electronegativity than calcium.
    d. Strontium, because strontium has a lower first ionization energy than calcium.

21. Neutral atoms of chlorine are bombarded by high-energy photons, causing the ejection of electrons from the various filled subshells. Electrons from which subshell would have the highest velocity after being ejected?
    a. 1s b. 2p c. 3p d. 3d

22. A photoelectron spectrum for which of the following atoms would show peaks at exactly three different binding energies?



(A) (B) (C) (D)

23. Most transition metals share a common oxidation state of +2. Which of the following best explains why?

    a. Transition metals all have a minimum of two unpaired electrons.
    b. Transition metals have unstable configurations and are very reactive.
    c. Transition metals tend to gain electrons when reacting with other elements.
    d. Transition metals will lose their outermost s-block electrons when forming bonds.

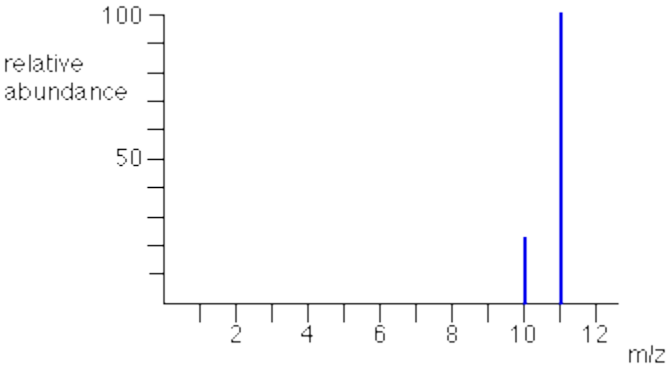
24. Some properties of scandium are determined by the electron arrangement within scandium. What is the ground state electron configuration of scandium?

    a. [Ar] $4s^2 3p^1$
    b. [Ar] $4s^2 3d^1$
    c. [Ar] $4s^2 4p^1$
    d. [Ar] $4s^2 4d^1$

25. Which of the following nuclei has 3 more neutrons than protons?

    a. ^{11}B
    b. ^{37}Cl
    c. ^{24}Mg
    d. ^{70}Ga

26. Given the mass spectrum and data for boron below, estimate the average atomic mass of boron.



Isotope	Peak Intensity
boron−10	23%
boron−11	100.%

    a. 10.20 amu
    b. 10.81 amu
    c. 10.98 amu
    d. 13.30 amu

27. Which of the following could have the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6$?

I. Mg^{2+} IV. Al^{3+}

II. Cl^- V. S^{2-}

III. K^+

a. II and V only
b. I, II, and IV only
c. II, III, and V only
d. I, IV, and V only

$1s^2 2s^2 2p^6 3s^2 3p^3$

28. Atoms of an element, X, have the electronic configuration shown above. The compound most likely formed with magnesium is:

a. Mg_2X b. MgX_2 c. MgX_3 d. Mg_3X_2

29. In which of the following groups are the three species isoelectronic?

a. S^{2-}, K^+, Ca^{2+}
b. Mg^{2+}, Ca^{2+}, Sr^{2+}
c. Sc, Ti, V^{2+}
d. Cs, Ba^{2+}, La^{3+}

30. Which of the following hydrogen electron transitions will result in the absorption of light with the highest energy?

a. n = 2 to n = 3
b. n = 3 to n = 2
c. n = 3 to n = 4
d. n = 4 to n = 3

31. Which is an electron configuration of a neon atom in the excited state?

a. $1s^2 2s^2 2p^4$
b. $1s^2 2s^2 2p^5$
c. $1s^2 2s^2 2p^4 3s^1$
d. $1s^2 2s^2 2p^5 3s^1$

32. Which of the following elements has the lowest first ionization energy?

a. arsenic
b. selenium
c. silicon
d. iodine

33. Consider the halogens chlorine and bromine. Which has a larger atomic radius and why?

a. Chlorine has a larger atomic radius because it has an increased number of principal energy levels.
b. Chlorine has a larger atomic radius because it has a higher effective nuclear charge.
c. Bromine has a larger atomic radius because it has an increased number of principal energy levels.
d. Bromine has a larger atomic radius because it has a higher effective nuclear charge.

34. Which of the following elements has the smallest atomic radius?

    a. nitrogen
    b. phosphorus
    c. oxygen
    d. sulfur

35. Which of the following elements has the highest first ionization energy?

    a. magnesium
    b. aluminum
    c. sulfur
    d. argon

36. Which of the following, sodium or magnesium, has a lower IE_1 and why?

    a. Sodium, because sodium has fewer principal energy levels than magnesium.
    b. Sodium, because sodium has a lower effective nuclear charge than magnesium.
    c. Magnesium, because of the repulsion of magnesium's paired 4s electrons.
    d. Magnesium, because magnesium has a greater effective nuclear charge than sodium.

37. Silicon has a larger atomic radius than phosphorus. Why?

    a. more principal energy levels
    b. fewer principal energy levels
    c. a larger nuclear charge
    d. a smaller nuclear charge

38. Which element forms an ion that is smaller than its atom?

    a. sodium
    b. oxygen
    c. chlorine
    d. phosphorus

39. Which family tends to form ions that are larger than their neutral atoms?

    a. transition metals
    b. alkaline metals
    c. halogens
    d. noble gases

40. Which of the following species has the greatest radius?

    a. S^{2-}
    b. Cl^{-}
    c. Ar
    d. K^{+}

41. Which of the following, Sr^{2+} and Sr, has a larger radius and why?

    a. Sr, because neutral Sr has more principal energy levels than Sr^{2+}.
    b. Sr, because neutral Sr has more protons than Sr^{2+}.
    c. Sr^{2+}, because Sr^{2+} has fewer principal energy levels than neutral Sr.
    d. Sr^{2+}, because Sr^{2+} has fewer electrons than neutral Sr.

42. Which of the following, lithium or gallium, has a lower electronegativity and why?

    a. Lithium, because lithium has fewer principal energy levels than gallium.
    b. Lithium, because lithium has a smaller effective nuclear charge than gallium.
    c. Gallium, because gallium has more principal energy levels than lithium.
    d. Gallium, because gallium has a larger effective nuclear charge than lithium.

43. Which of the following elements has the highest first ionization energy?

    a. Na
    b. K
    c. P
    d. S

44. Which of the following is an accurate description of the electronegativities of elements in the periodic table?

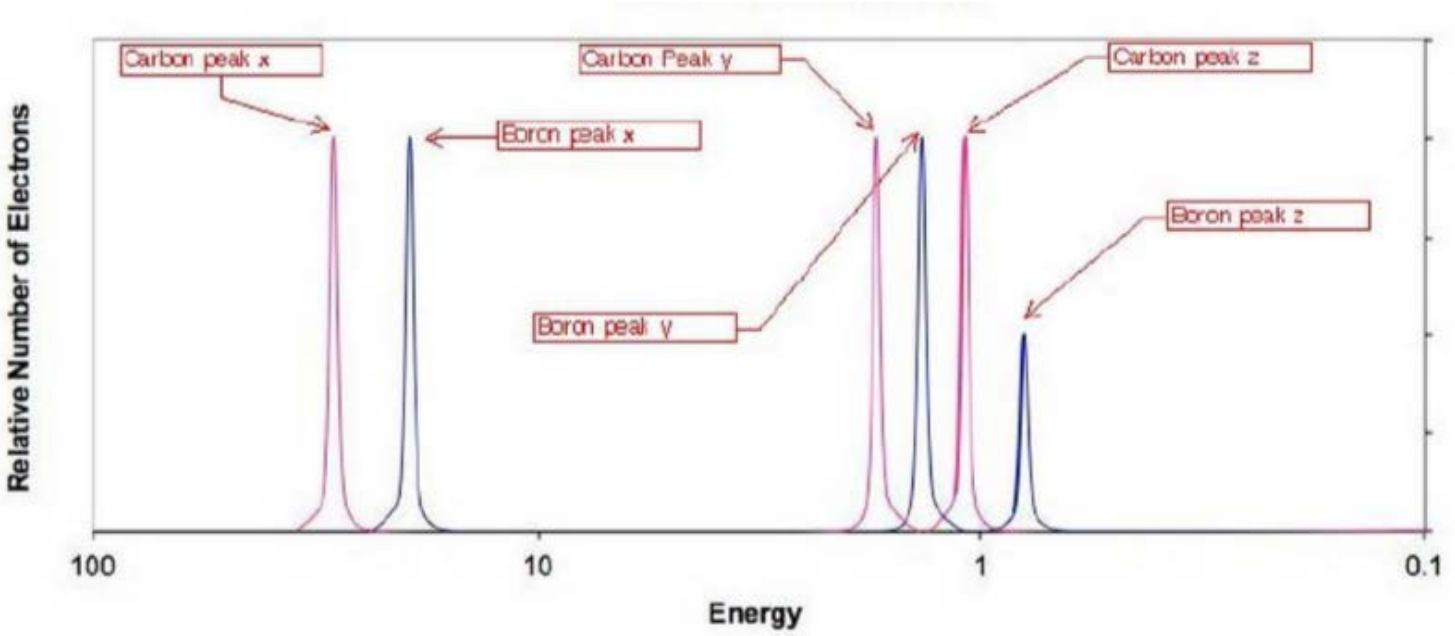
    a. The electronegativity of bromine is greater than that of chlorine.
    b. The electronegativity of argon is greater than that of chlorine.
    c. The electronegativity of calcium is greater than that of potassium.
    d. The electronegativity of aluminum is greater than that of sulfur.

45. Which of the following species is NOT isoelectronic to Br^{-}?

    a. Se^{2-}
    b. Rb^{+}
    c. Kr
    d. K^{+}

Free Response Practice #1 (10 points)

1. Consider the PES spectra shown below, that superimposes the simulated PES spectra for elemental boron and elemental carbon on the same plot. NOTE: x, y, and z are simply labels and do NOT refer to p_x, p_y, and p_z orbitals.



a. Suggest a reason why the boron and carbon peaks have been paired together in three groups labeled x, y, and z. (1 point)

b. Explain why the carbon 'x peak' is at a higher energy than the boron 'x peak'. (2 points)

c. Explain why the boron 'z peak' is half the height of the carbon 'z peak'. (1 point)

d. If one were to superimpose a third PES plat on the same axes for elemental nitrogen;

   i. Relative to carbon's 'x peak', where would nitrogen's 'x peak' appear on the x-axis. Explain AND draw and label the peak on the plot above. (2 points)

   ii. Relative to carbon's 'z peak', what would the height of nitrogen's 'z peak' be? Explain AND draw and label the peak on the plot above. (2 points)

e. Identify the electrons that are associated with each of the following peaks on the boron plot: (2 points)

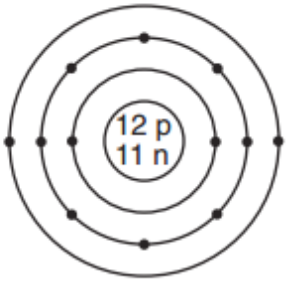
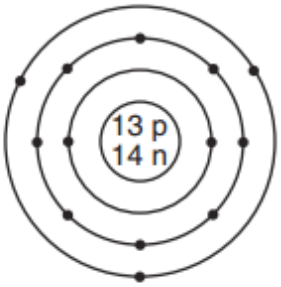
   i. x peak ii. y peak iii. z peak

Free Response Practice #2 (10 points)

2. Answer the following questions related to the properties of magnesium and aluminum.

Atomic Diagrams of Magnesium and Aluminum

Key
• = electron

Element	Lewis Electron-Dot Diagram	Electron-Shell Diagram
magnesium		
aluminum		

a. Write the complete electron configuration for a neutral atom of magnesium in the ground state. (1 point)

b. Write the complete orbital diagram for a neutral atom of aluminum in the ground state. (1 point)

c. The table below represents the first ionization energy for the elements in Period 3. The missing first ionization energy values are 496 kJ/mol, 578 kJ/mol, 738 kJ/mol, and 789 kJ/mol. Use these values and your understanding of the trend of first ionization energy to complete the table below. (2 points)

Element	Atomic Number	Symbol	First Ionization Energy (kJ/mol)
Sodium	11	Na	
Magnesium	12	Mg	
Aluminum	13	Al	
Silicon	14	Si	
Phosphorus	15	P	1012
Sulfur	16	S	1000
Chlorine	17	Cl	1251
Argon	18	Ar	1521

d. Why is the first ionization energy of sulfur lower than that of phosphorus? Explain using Coulomb's Law. (2 points)

e. Consider the table of successive ionization energies shown below.

Ionization Energies (kJ/mol)							
1st IE	2nd IE	3rd IE	4th IE	5th IE	6th IE	7th IE	8th IE
578	1,820	2,750	11,600	14,800	18,400	23,300	27,500

a. Which element from Period 3 does the table above represent? Explain. (2 points)

b. Write the electron configuration for the most commonly formed ion of the element chosen in part (e). (1 point)

c. Identify one other neutral atom or ion that is isoelectronic with the most commonly formed ion of this element. (1 point)

Free Response Practice #3 (6 points)

3. Answer the following questions related to aluminum and one of its compounds.
    a. Consider the two chemical species Al and Al^{3+}.
        i. Write the electron configuration of each species. (2 points)

        ii. Why is the radius of the Al atom larger than the radius of the Al^{3+} ion? Explain using the principles of atomic structure. (2 points)

    b. The Al^{3+} ion is isoelectronic with the Ne atom. From which species, Al^{3+} or Ne, is it easier to remove an electron? Justify with Coulomb's Law. (2 points)

Free Response Practice #4 (7 points)

4. Suppose that a stable element with atomic number 119, symbol Q, has been discovered.

   a. Write the ground-state electron configuration for Q, showing only the valence-shell electrons. (1 point)

   b. Would Q be a metal or a non-metal? Explain in terms of atomic structure. (2 points)

   c. On the basis of periodic trends, would Q have the largest atomic radius in its group or would it have the smallest? Explain in terms of electronic structure. (2 points)

   d. What would be the most likely oxidation state of the Q ion in a stable ionic compounds? (1 point)

   e. Assume that Q reacts to form a carbonate compound. Write the formula for the compound formed between Q and the carbondate ion, CO_3^{2-}. (1 point)

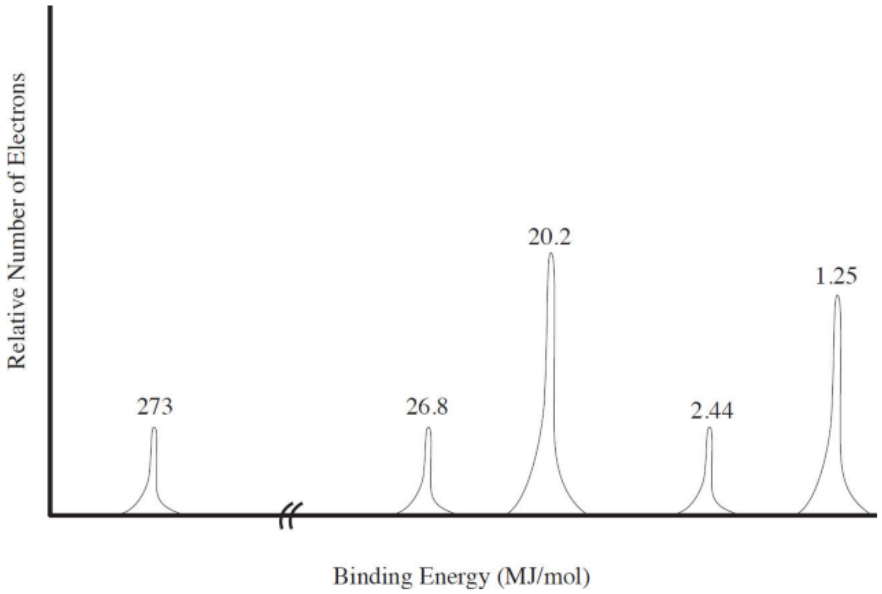
Free Response Practice #5 (6 points)

5. Explain the following observations using Coulomb's Law. In each part, your answer must include specific information about both substances.
    a. It requires less energy to remove the first electron from magnesium than the second electron (i.e. $IE_1 < IE_2$). (2 points)

    b. The electronegativity of sulfur is less than the electronegativity of chlorine. (2 points)

    c. The Ca^{2+} and the Cl^- ions are isoelectronic, but Cl^- has a larger radius. (2 points)

Free Response Practice #6 (Cracking the AP Chem Exam, Practice FR #4, edited, 7 points)



1. The above PES belongs to a neutral chlorine atom.
    a. What energy of light, in MJ/mol, would be required to eject a 3*s* electron from chlorine? (1 point)

    b. What wavelength of light, in m, would be required to eject the same 3*s* electron? (2 points)

A second PES is run, this time for a sample of chloride ions.

c. For the PES of a chloride ion, how would the following variables compare to the peaks on the PES above? Justify your answers.

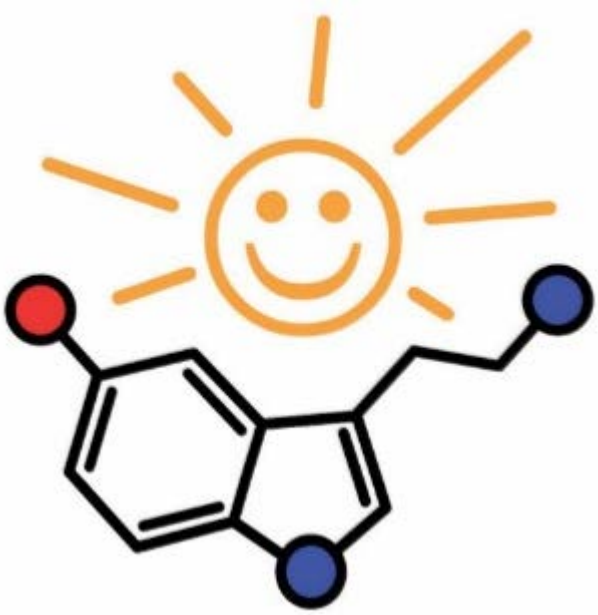
    i. Number of peaks (1 point)

    ii. Height of peaks (1 point)

d. Draw the orbital diagram for a chloride ion. (1 point)

e. Identify the noble gas which is isoelectronic to a chloride ion. (1 point)

AP Chemistry FTW!



Unit 11:

Bonding & IMFs

Unit 11 Objectives

BIG IDEA 1 - The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identity in chemical reactions.

- Enduring Understanding 1.E: Atoms are conserved in physical and chemical processes.

BIG IDEA 2 - Chemical and physical properties of materials can be explained by the structure and arrangement of atoms, ions, molecules and the forces between them.

- Enduring Understanding 2.A: Matter can be described by its physical properties. The physical properties of a substance generally depend on the spacing between the particles (atoms, molecules, ions) that make up the substance and the forces of attraction among them.
- Enduring Understanding 2.B: Forces of attraction between particles (including the noble gases and also different parts of large molecules) are important in determining many macroscopic properties of a substance, including how observable physical state changes with temperature.
- Enduring Understanding 2.C: The strong electrostatic forces of attraction holding atoms together in a unit are called chemical bonds.
- Enduring Understanding 2.D: The type of bonding in the solid state can be induced from the properties of the solid state.

BIG IDEA 5 - The laws of thermodynamics describe the essential role of energy and explain and predict the direction of changes in matter.

- Enduring Understanding 5.B: Energy is neither created nor destroyed, but only transformed from one form to another.
- Enduring Understanding 5.C: Breaking bonds requires energy, and making bonds releases energy.
- Enduring Understanding 5.D: Electrostatic forces exist between molecules as well as between atoms or ions, and breaking the resultant intermolecular interactions requires energy.

Bonding Day 1: Can you hold it together??

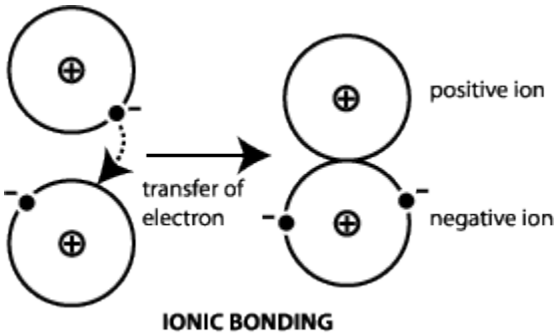
Bonding: The attractive forces that hold groups of atoms together are called chemical _____.

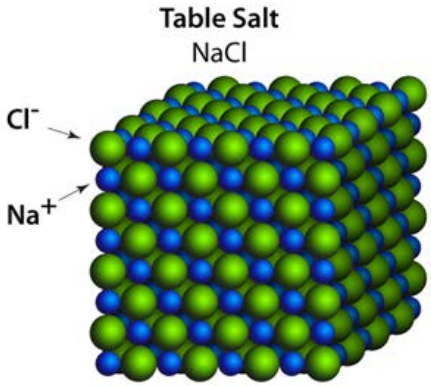
- ➔ Bonds tend to form so that each atom, by gaining, losing, or sharing electrons has **eight** electrons in its valence level (this is called the _____ rule).
- ➔ The goal of bonding is to achieve the _____ possible energy state.

Types of Bonds		
Ionic	**Covalent**	**Metallic**

Ionic Bonding: Metal + Nonmetal(s) (atoms with very _____ electronegativities)

- Electrons are _____ from the _____ to the _____.
- Electrical attraction between a _____ (metal) and an _____ (nonmetal).
- Ionic compounds form a solid, regular array of cations and anions called a _____.
- **Lattice energy:** how much energy it takes to _____ apart a solid ionic compound.





Ionic Bond Properties

- _____ melting points and boiling points.
- hard and brittle
- conduct _____ when liquid or aqueous (dissolved in water) but not when solid

Lewis Dot Structures of Binary Ionic Compounds

Metal ions are always _____! Non-metal ions are always _____!

1) magnesium fluoride	2) gallium chloride

Key Formulas and Relationships

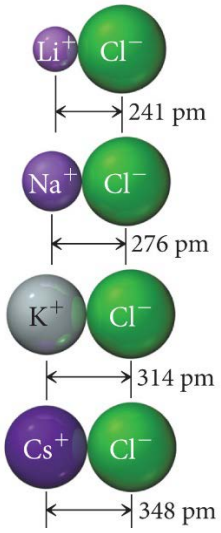
When answering questions about ionic bond strength, justify your response using Coulomb's Law:

$$Lattice\ Energy = k(\frac{Q_1 Q_2}{d})$$

Use Coulomb's Law to justify melting point, solubility, and lattice energy differences between two ionic compounds.

➔ The more highly _____ the ions OR the _____ the ions, the **GREATER** the attraction!

Lattice energy: energy released when the solid crystal forms from separate ions in the gas phase



- Directly dependent on size of charges
- Inversely dependent on distance between ions
- Ion charge is generally _____ important than ion size

Greater lattice energy = _____ energy required to separate ions

→ _____ ionic bond

→ ____ melting point

→ ____ solubility (ions must separate/dissociate from one another and attach to water to dissolve)

Metal Chloride	Lattice Energy kJ/mol
LiCl	−834
NaCl	−788
KCl	−701
CsCl	−657

Let's Practice!

1. Arrange the following ionic compounds NaF, MgF_2, MgO, KF
   a. in order of increasing lattice energy:

   b. in order of increasing melting point:

   c. in order of increasing solubility:

2. Would the lattice energy of lithium fluoride be larger or smaller than the lattice energy of potassium bromide. Explain in terms of Coulomb's Law.

Percent Ionic Character:

- The greater the difference in _____ between two bonded atoms, the greater the ionic character of the bond.
- The more _____ in electronegativity, the greater the covalent character of the bond.

Dipole moment: a measure of bond polarity; ____ dipole moment means ____ ionic character!

➔ Represented by an arrow pointing in the direction of greater electron density 

TABLE 9.2 Dipole Moments of Several Molecules in the Gas Phase

Molecule	ΔEN	Dipole Moment (D)
Cl_2	0	0
ClF	1.0	0.88
HF	1.9	1.82
LiF	3.0	6.33

Let's Practice! Given the dipole moments listed in the table above, arrange the four molecules:

a. in order of increasing ionic character:

b. in order of increasing covalent character:

Multiple Choice Practice

1. A crystalline solid has a melting point of 502°C, and it conducts electricity in an aqueous solution, but not while solid. Which of the following is most likely to be the identity of the substance?

   a. $I_2(s)$ c. $C_{12}H_{22}O_{11}(s)$

   b. LiCl(s) d. Ni(s)

2. Which of the pairs of ions listed below will form a salt with the greatest melting point?

   a. K^+ and Cl^- c. Ca^{2+} and S^{2-}

   b. Rb^+ and Cl^- d. Sr^{2+} and S^{2-}

3. Which of the following substances has the greatest ionic character?

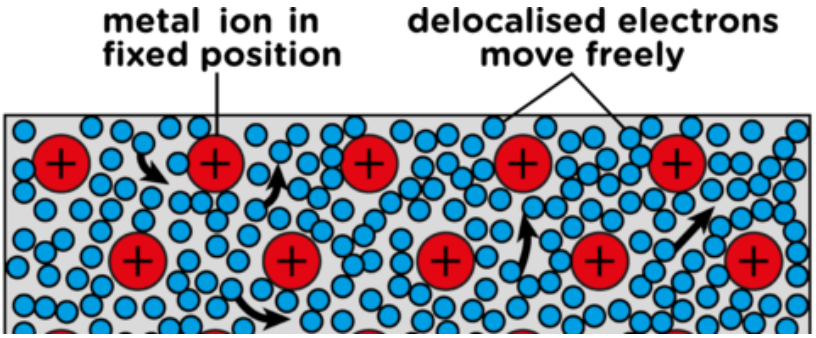
   a. XeF_2 c. AsP_3

   b. $GaBr_2$ d. PCl_5

Metallic Bonding and Alloys

Metallic bonds: results from the attraction between metals atoms and the surrounding _____.

- The valence electrons of metal atoms are not tightly held, which means its valence electrons are _____ (not stuck in one place) and _____ (free to move around).
- The sea of electrons occupies the interstitial spaces (the spaces between the metal cores).
- Because each metal atom allows its electrons to roam freely, these atoms become positively charged cations. The attraction between the _____ electrons and _____ metal cores is a metallic bond.



Metallic bond properties:

1. Highly conductive of heat and electricity (because of mobile electrons).
2. Lustrous (_____) because it reflects incoming light photons (because of the mobile electrons.)
3. The structure and uniform bonding in all directions of the metal allow the atoms to slide past each other without breaking, leading to:
   a. Malleability: ability of a substance to be hammered or beaten into thin sheets.
   b. Ductility: ability of a substance to be drawn or pulled through a small opening to produce a _____.

Let's Practice! Explain the following in terms of structure and/or bonding.

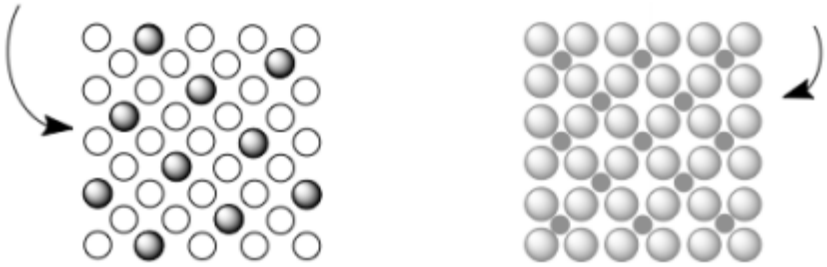
1. Solid K conducts a strong electric current, whereas solid KNO_3 does not.

2. Which of the following solids will conduct an electrical current? Explain. Mg or NaCl or CO_2

Alloys: similar in structure to pure metal solids, but contain _____ than one type of element.

There are two types of alloys that are AP tested!

Substitutional Alloys	Interstitial Alloys
Form between atoms of _____ size, where one atom substitutes for the other in the lattice. • Similar properties to component atoms • _____ malleable and ductile	Form between atoms of _____ size, where the smaller atoms fill the interstitial spaces (lattice holes) between the larger atoms. • Properties change!! • More _____ • _____ malleable and ductile

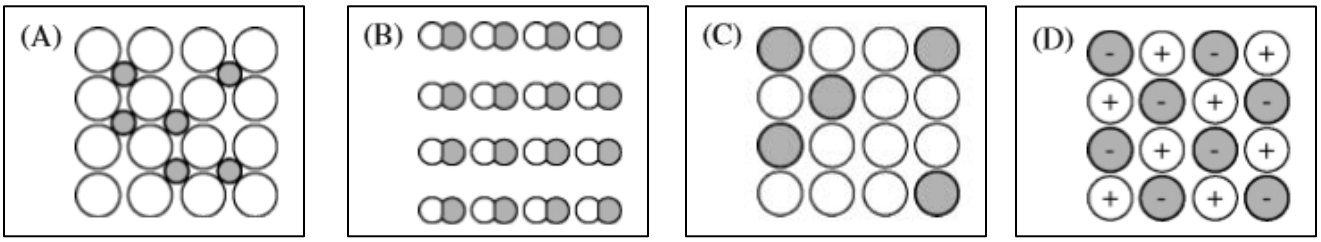


1. Alloys typically retain a sea of mobile electrons and so remain _____.
2. In some cases, alloy formation alters the chemistry of the surface. An example is the formation of a chemically inert _____ layer in stainless steel.

Let's Practice!

1. To increase its strength and hardness, gold alloys are created by combining it with other metals such as nickel and palladium (Pd). When comparing a Au/Ni alloy with a Au/Pd alloy, both made with the same mole fraction of gold, the Au/Ni alloy is measurably harder than the Au/Pd alloy. Which of the following statements best explains why?
    a. Ni has only one common oxidation state, but Pd has two.
    b. Pd has a lower melting point than Au, but Ni has a higher melting point.
    c. Ni atoms are smaller than Pd atoms, and so they interfere more with the displacement of atoms in the alloy.
    d. Ni atoms are less polarizable than either Au or Pd atoms, and so Ni has weaker interparticle forces.

2. Which diagram shown below would be the best representation of steel: an alloy of iron and carbon?



Covalent Bonds

Covalent (Molecular) Bonding: two nonmetal atoms _____ electrons to fill the valence level of both atoms.

- Occurs between elements with _____ electronegativities, high effective nuclear charges (Z_{eff}) and small radii, so they can attract and hold each other's electrons to make shared pairs of electrons.
- The smallest group of elements held together by a covalent bond is called a _____.
- Atoms can make single, double, or triple bonds depending on whether they share one, two, or three pairs of electrons respectively.
    - **Multiple bonds are most often formed by C, N, O, P and S atoms,** aka _____.

Covalent Bond Properties

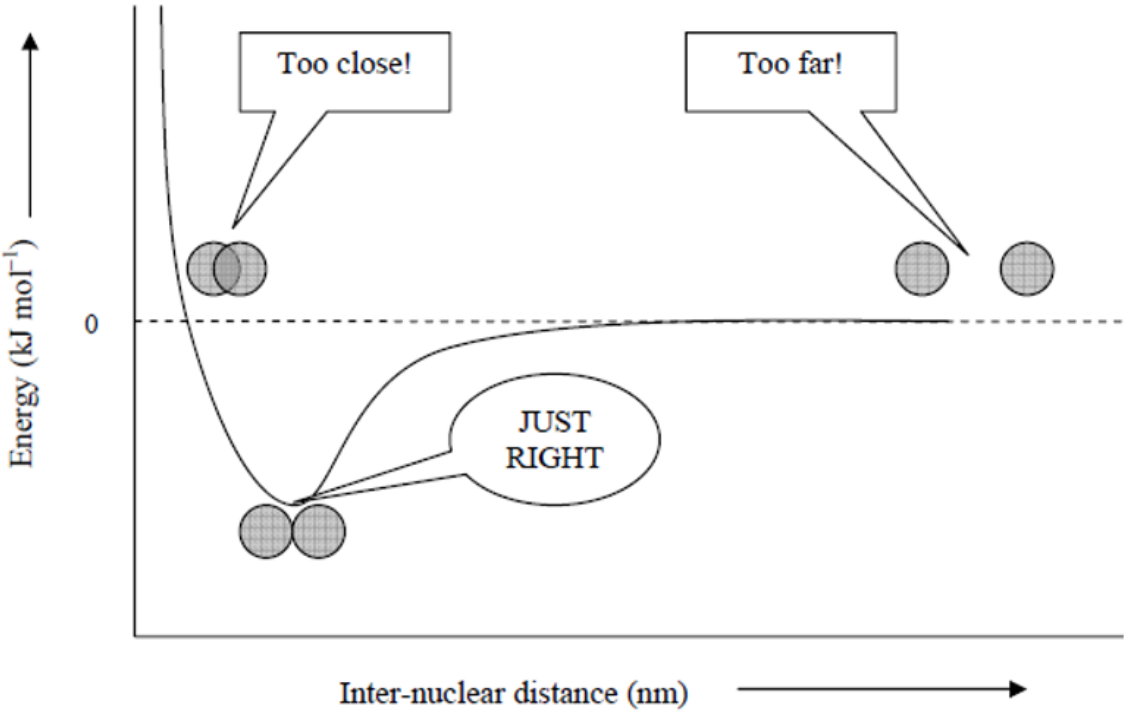
- _____ melting points and boiling points.
- Typically do _____ conduct electrical current because they lack mobile charges (EXCEPT strong acids!)

Coulomb's Law tells us:

- The negative electrons of one atom and the positive nucleus of another atom _____ each other.
- If the nuclei of two atoms get too close together, their like charges _____ each other.

Bond length: the distance two covalently bonded atoms at their _____ potential energy. It is a balance between opposing forces:

- Attractive electrostatic forces between the nucleus of one atom and the electrons of another
- Repulsive forces between the two positively charged nuclei



Bond Order: the _____ of chemical bonds between a pair of atoms; indicates the stability of a bond.

Bond Type	Bond Order	Bond Length	Bond Strength
Single bond	1	longer	weaker
Double bond	2	medium	medium
Triple bond	3	shorter	stronger

Multiple bonds increase the electron density between two nuclei

- Decreases repulsions between the two nuclei
- Added electrons ____ attraction between nuclei and electron density → multiple bonds ____ bond strength!
- Nuclei can move closer together → multiple bonds ____ bond length!

Directions: On the left, show the neutral, separate atoms using Lewis valence electron dot structures. On the right, depict the bonding atoms sharing electrons.

2 atoms of F	**F_2**
2 atoms of O	**O_2**
2 atoms of N	**N_2**

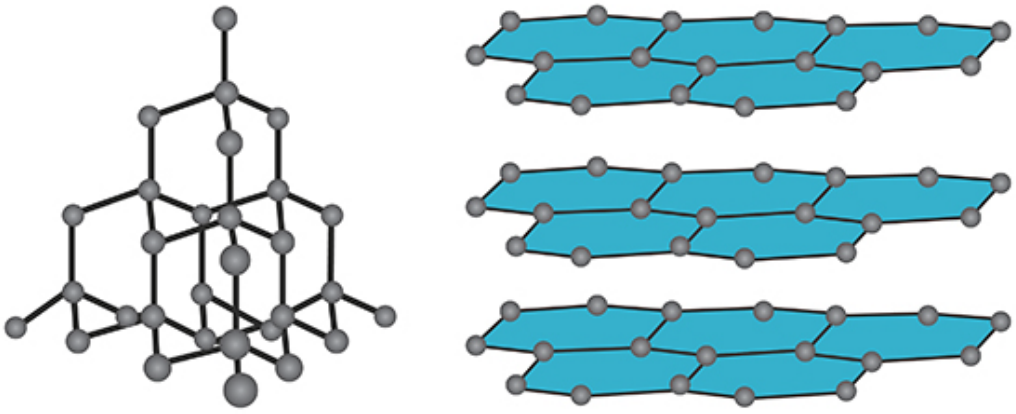
1. Explain: why does it require more energy to break the bond between O_2 than F_2?

2. Rank the following in order of increasing bond length: O_2, F_2, N_2, H_2

Network Covalent Solids

Network Covalent Solid: a crystalline structure is formed by non-metals covalently bonded together into 2-D (sheets) or 3-D networks.

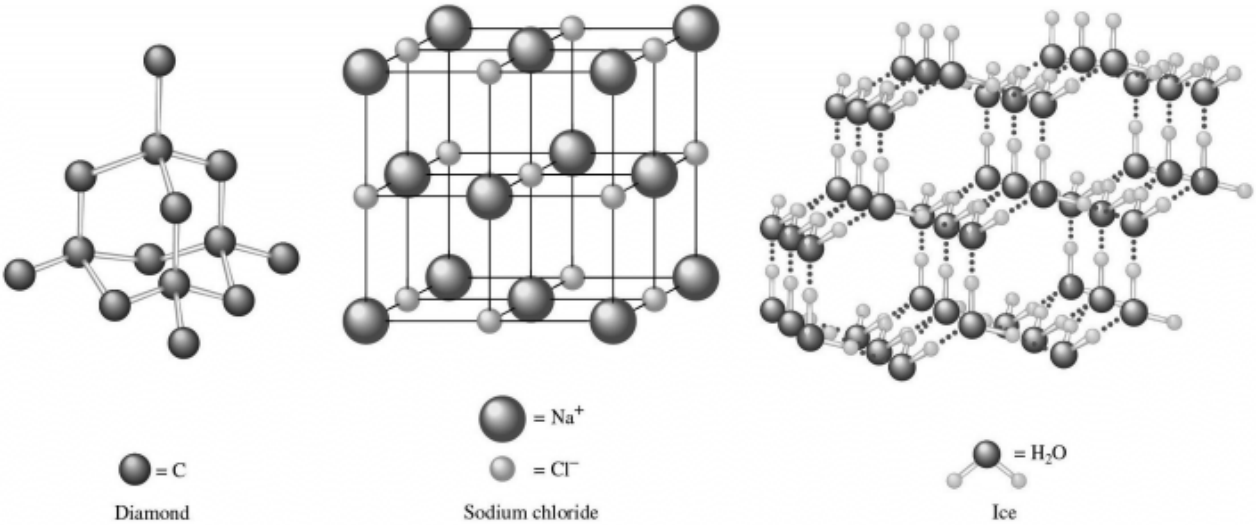
- A perfect single crystal of a network covalent solid is a single, giant molecule!
- VERY _____ melting and boiling points
- Very rigid and hard
- Chemically _____ (non-reactive); rarely dissolve in water



Diamond Graphite

Types of Crystalline Solids

Network Covalent Solids	**Ionic Solids**	**Molecular Solids**
e.g. giant molecules carbon (diamond, graphite)	e.g. salts like NaCl	e.g. protein crystals, sucrose



Diamond Sodium chloride Ice

In summary:

Types of Chemical Bonds				
	Made of	**One unit is called**	**What holds them together?**	**Characteristic Properties**
Ionic	metal + non-metal(s)	formula unit	• electrostatic attraction: between cation (+) and anion (−)	• High melting and boiling points
Covalent	non-metal + non-metal	molecule	• shared electrons	• Low melting and boiling points (unless a network covalent solid such as diamond)
Metallic	metals only	metal 😉	• electrostatic attraction: between "sea" of mobile or delocalized electrons and positive metal ions	• Good conductors of heat and electricity (because of mobile electrons)

Let's Practice!

1. When phosphorus and sulfur bond, the reaction involves a:

   a. creation of electrons
   b. sharing of electrons
   c. transfer of electrons from P to S
   d. transfer of electrons from S to P

2. Which bond is the strongest?

   a. C = C b. C – H c. C = N d. N ≡ N

3. Which of the following substances is an electrolyte when dissolved in water?

   a. CH_3CH_2OH
   b. SiF_4
   c. HBr
   d. OF_2

4. When calcium and fluorine bond:

   a. two F atoms each give 7 electrons to Ca.
   b. Ca loses 2 electrons, 1 to each fluorine.
   c. two F atoms each give 1 electron to Ca.
   d. Ca shares electrons with F, forming 2 Ca – F bonds.

5. Which of the following has the lowest ionic character?

   a. KCl b. $(NH_4)_2CO_3$ c. SO_2 d. MgF_2

6. Which of the following is NOT a property of an ionic compound?

   a. hard
   b. brittle
   c. melts readily at room temperature
   d. conducts electricity when dissolved in water

7. Which of the following substances would be predicted to have the highest melting point?

   a. H_2S b. AlF_3 c. P_2O_5 d. H_2

8. Which of the following is NOT a property of a metallic compound?

   a. malleable
   b. ductile
   c. mobile electrons easily conduct heat and electricity
   d. held together by electrostatic attraction between cations and anions

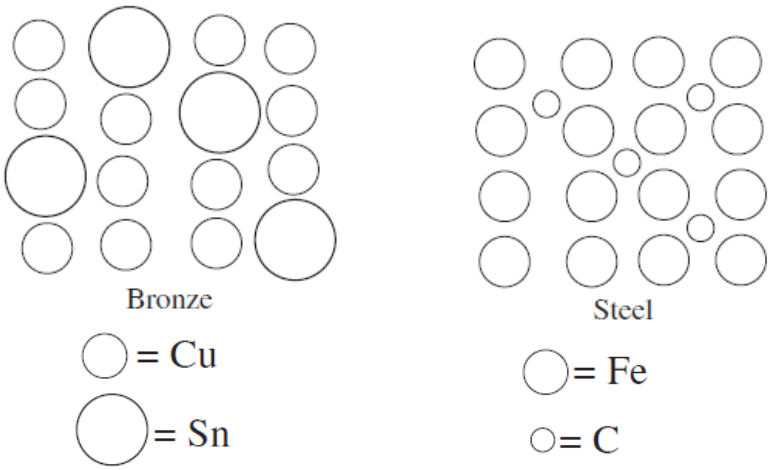
9. When iron and oxygen bond, the reaction involves a:

   a. creation of electrons
   b. sharing of electrons
   c. transfer of electrons from Fe to O
   d. transfer of electrons from O to Fe

10. What substance will NOT conduct electricity when dissolved in water?

    a. $C_6H_{12}O_6$
    b. $(NH_4)_2CO_3$
    c. HCl
    d. AlF_3

11. Two alloys are shown in the diagrams below – bronze and steel. Which of the following correctly describes the malleability of both alloys compared to their primary metals?



    a. Bronze's malleability would be comparable to that of copper, but steel's malleability would be significantly lower than that of iron.
    b. Bronze's malleability would be significantly higher than that of copper, but steel's malleability would be comparable to that of iron.
    c. Both bronze and steel would have malleability values similar to those of their primary metals.
    d. Both bronze and steel would have malleability values lower than those of their primary metals.

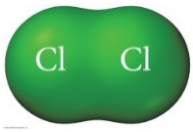
12. Which compound $CaCl_2$ or CaO, would you expect to have a higher melting point? Why?

    a. $CaCl_2$, because there are more ions per lattice unit.
    b. $CaCl_2$, because a chlorine ion is smaller than an oxygen ion.
    c. CaO, because the charge of an oxygen ion exceeds that of a chlorine ion
    d. CaO, because the common charges of calcium and oxygen atoms are identical in magnitude

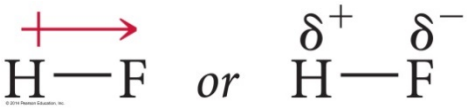
Non-Polar Covalent, Polar Covalent, or Ionic Bonding?

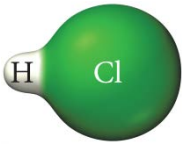
The attraction or "pull" on the bonded electron pair (i.e. _____) determines bond polarity.

1. **Non-polar** covalent bond: bonding electrons are shared _____ by the bonded atoms.
   - Electronegativity difference between atoms (ΔEN) < _____.
   - Examples:
     i. Diatomic molecules (Br_2, I_2, N_2, etc)
     ii. Any _____ bond.

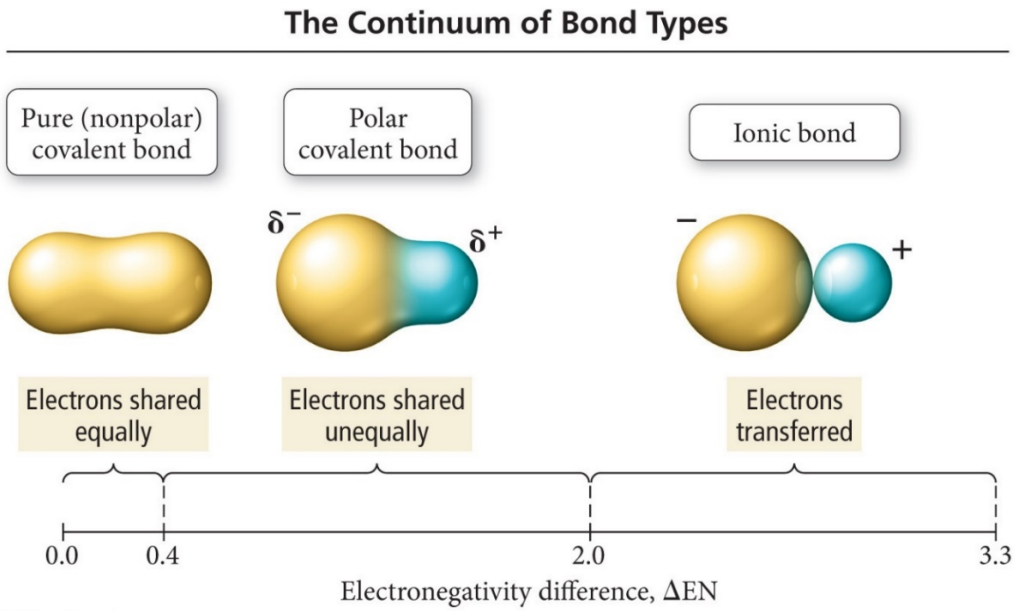


2. **Polar** covalent bond: bonded atoms have an _____ attraction for the shared electrons.
   - _____< ΔEN < _____
   - The atom with the greatest _____ has a greater attraction for the shared electrons, so they claim a greater amount of electron density.
   - The uneven electron density creates a _____: a partial _____ charge on the atom with higher electron density and a partial _____ charge on the atom with lower electron density.
   - An _____ is used to represent the dipole (sometimes called a dipole moment): the arrow points towards the _____ pole (i.e. the most electronegative atom) and has a crossed tail at the _____ pole (least electronegative atom).





3. **Ionic** bond: a bond where the electronegativity difference between the two atoms is so extreme that one atom takes custody of all the contested electrons! (ΔEN > _____)



Most chemical bonds are somewhere between purely ionic and purely covalent!

All Together Now ☺

Remember, nature is striving for a LOWER ENERGY STATE!

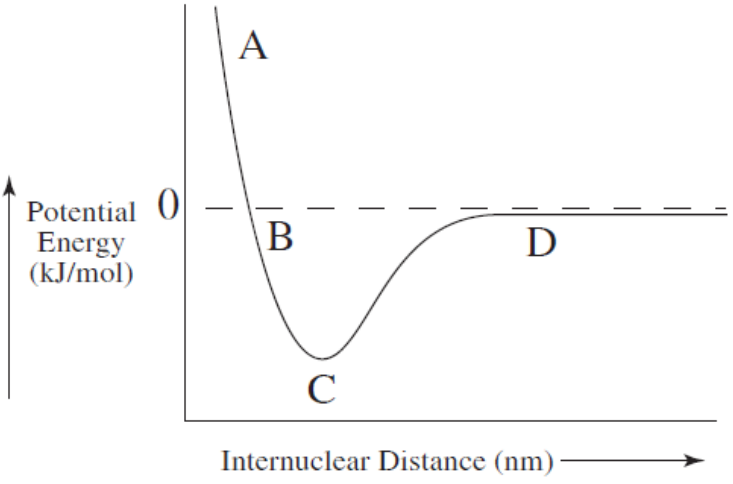
Types of Chemical Bonds			
	What holds them together?	**Characteristic Properties**	**Conductivity**
Ionic	Strong electrostatic attraction between ions (lattice energy)	• Typically high melting and boiling points; • Usually found in the solid state because the electrostatic attraction is SO strong	• Conductors of electricity **only** in (aq) or (l) states (when ions are mobile) • Electrolytes when dissolved in water
Covalent	Electrons are shared by nuclei; **however**, sharing is hardly ever equal!	• Can be solids, liquids or gases (depending on IMFs); • Have low melting and boiling points	• Poor conductors of electricity (no mobile charges) • Not electrolytes when dissolved in water
Metallic	Attraction between "sea" of mobile or delocalized electrons and positive metal ions	• Solids with a crystalline structure at room temp; • Range of melting points, usually depending on number of valence electrons	• Excellent conductors of electricity since electrons in the "sea" are free to move

More Practice!

1. An unknown substance is found to have a high melting point. In addition, it is a poor conductor of electricity and does not dissolve in water. The substance most likely contains

   a. ionic bonding
   b. metallic bonding
   c. non-polar covalent bonding
   d. covalent network bonding

2. The graph below shows the amount of potential energy between two hydrogen atoms as the distance between them changes. At which point in the graph would a molecule of H_2 be the most stable?



   a. Point A
   b. Point B
   c. Point C
   d. Point D

Covalent Lewis Dot Structures

Covalent Lewis Dot Structures: formulas used to model what atoms look like in a compound that contains atoms that are covalently bonded together.

- Non-metals will share electrons to get ____ valence electrons and be stable.

Helpful Hints for Drawing Lewis Dot Structures

1. H is always a terminal atom → ALWAYS connected to only _____ other atom.
2. _____ electronegativity is central atom in molecule.
3. If drawing the Lewis structure for a polyatomic ion,
   a. For positive ions, _____ electron(s) from the central atom.
   b. For negative ions, _____ electron(s) to the central atom.
   c. Enclose the dot structure in square _____ and include the ion's charge outside the brackets.
4. Not all elements can form double or triple bonds: **only C, N, O, P, and S!** (Think CNOP-S)
5. For molecules with more than one central atom, use the formula to help you decide how to draw the structure.
6. The total number of valence electrons in your Lewis structure MUST equal the _____ of the valence electrons of all of the elements added together (add or subtract electrons as needed for polyatomic ions).
7. For *most* covalent compounds, you can determine how many bonds each atom will form by looking at the number of unpaired electrons in their Lewis dot structure.
   a. Paired electrons do _____ form bonds.
   b. Single electrons _____ form bonds!

H	**Be**	**B**	**C**	**N**	**O**	**F**
# bonds? ____	# bonds? ____	# bonds? ____	# bonds? ____	# bonds? ____	# bonds? ____	# bonds? ____

Examples:

CH_4	CO_2	CH_2O

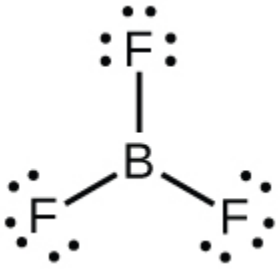
NH_2^-	NH_3	NH_4^+
OH^-	H_2O	H_3O^+

N_2H_2	CH_2CH_2
CH_3CHCH_2	CH_3CH_2COOH

But wait!!! Exceptions to the octet rule

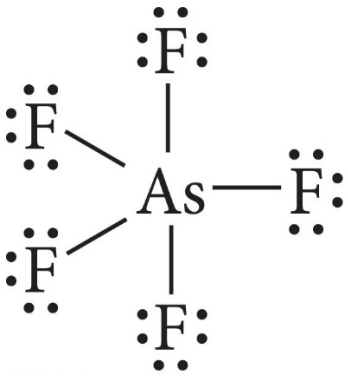
1. Elements that will have _____ than 8 valence electrons and are stable.

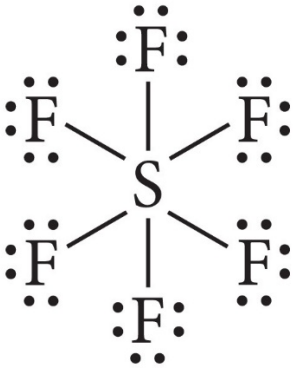
- **a.** _____, 2 electrons (___ bond)
- **b.** _____, 4 electrons (___ bonds)
- **c.** _____, 6 electrons (___ bonds)



2. Elements that will have _____ than 8 valence electrons and are stable.

- **a.** Elements in period (row) _____ through _____ can often expand their octet and can form more than 4 bonds (can have up to 12 electrons, 6 bonds)
  1. This is only possible between periods 3 through 7 because they can hold electrons in their empty _____.
  2. If you are unsure where to put extra lone pairs, check to see if the <u>central atom</u> can have an expanded octet (check to see if the element is in periods 3 through 7)





Challenge: Lewis Dot Practice, including exceptions to the octet rule. ☺

$RnCl_2$	BeH_2
SF_6	BF_3
$XeCl_4$	IF_7

Formal Charge

Formal Charge: a way to identify the _____ Lewis dot structure when more than one valid dot structure exists

- Formal charges are hypothetical charges assigned to each _____ in the dot structure

Formal Charge = # of valence electrons − # non-bonding electrons (lone) − ½ # bonding electrons

You do NOT need to show work for formal charge calculations!!! 😊

Formal Charge Rules (Which dot structure is best?)

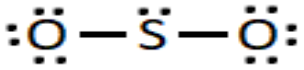
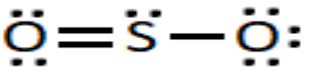
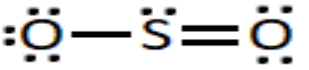
1) Small (or even better, _____) formal charges are more stable.
2) ____ formal charge on the _____ electronegative atoms.
3) ____ formal charge on the _____ electronegative atoms.
4) Sum of all formal charges must _____ the charge of the molecule.

Hint: Formal charge = _____ when the atom forms the number of bonds you would predict based on its Lewis structure.

Formal Charge = 0					
H	B	C	N	O	F

Let's Practice!

The compound SO_2 can be drawn with multiple valid Lewis dot structures:

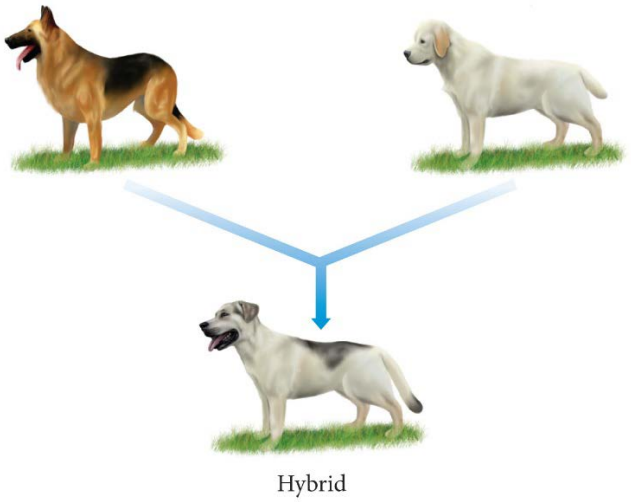
			
# of valance e^-			
− # of nonbonding e^-			
− ½ (# of bonding e^-)			
Formal Charge			

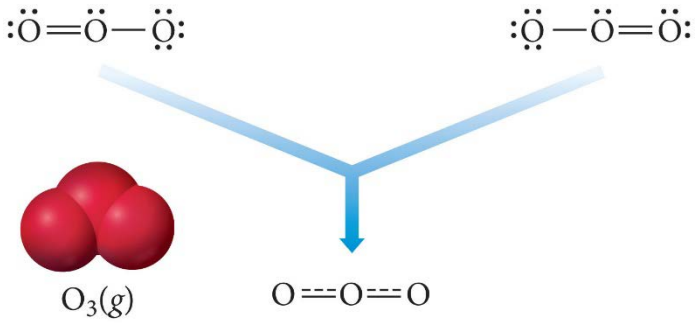
1. Which structure(s) best represents a molecule of SO_2? Justify your answer in terms of formal charge.

Resonance Structures

Resonance structures: when _____ or more Lewis structures can validly represent a molecule (or ion)

- Each resonance structure _____ to the real (observed) structure.
- The atoms of the molecule (or ion) stay in the _____ relative position: only the distribution of electrons is different!
- The actual structure, the resonance hybrid, is _____ between the two or more resonance structures.
    - All possible dot structures contribute to the real structure, BUT more stable ones (____ formal charge) contribute more.



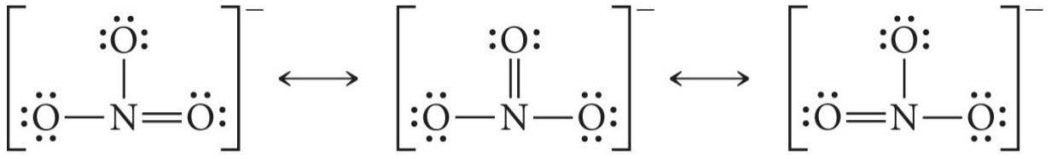


Important Notes about Resonance Structures:

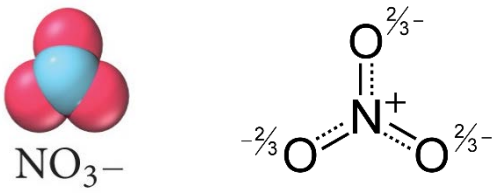
- Bonds are more equivalent to a "bond and a half" or a "bond and a third" in terms of length and strength, and are represented by one full line and one dashed line.
- Double edged arrows are used to indicate resonance.
- Resonance structures often occur in compounds with a _____ or _____ bond.
- Resonance _____ the stability of the molecule!

Example: Nitrate (NO_3^-)

Resonance Structures:



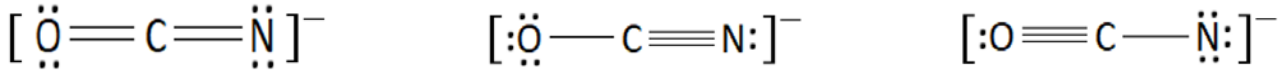
Resonance Hybrid (Real, Observed Structure):



Let's Practice!

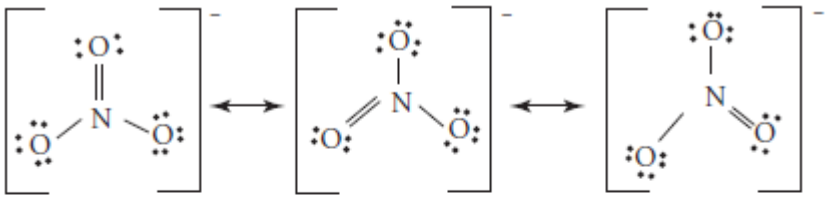
1. The ion OCN^- has three resonance structures.

   a. Identify the formal charge on each atom, for each resonance structure.

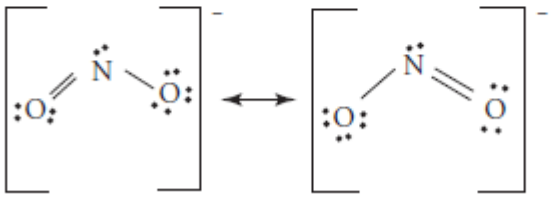


   b. Which resonance form is likely to contribute the most to the correct structure of OCN^- and why? Justify your answer in terms of formal charge.

2. Lewis diagrams for the nitrate and nitrite ions are shown below. Choose the statement that correctly describes the relationship between the two ions in terms of bond length and bond energy.



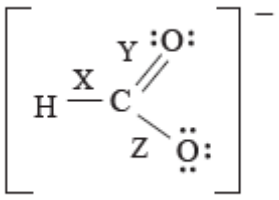
Nitrate



Nitrite

   a. Nitrite has longer and stronger bonds than nitrate.
   b. Nitrite has longer and weaker bonds than nitrate.
   c. Nitrite has shorter and stronger bonds than nitrate.
   d. Nitrite has shorter and weaker bonds than nitrate.

3. The formate ion, HCO_2^-, is best represented by the Lewis diagram below. Each bond is labeled with a different letter.

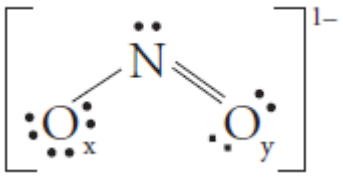


What is the bond order for each bond?

	X	Y	Z
(A)	1	1	2
(B)	2	2	1
(C)	1	1.5	1.5
(D)	1.33	1.33	1.33

4. Atoms of which element are most likely to form a structure with the formula XF_6 (where X is one of the four atoms)?

a. Carbon
b. Neon
c. Nitrogen
d. Sulfur



5. One of the resonance structures for the nitrite ion is shown above. What is the formal charge on each atom?

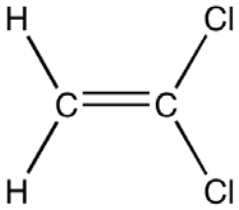
	O_x	N	O_y
(A)	–1	+1	–1
(B)	+1	–1	0
(C)	0	0	–1
(D)	–1	0	0

Isomers

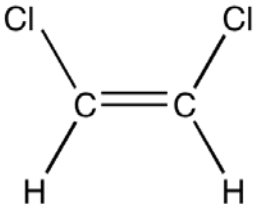
Isomers: molecules with the _____ molecular formula but _____ spatial arrangement of atoms

- Same _____ of atoms of each element
- Different _____ of their atoms in space

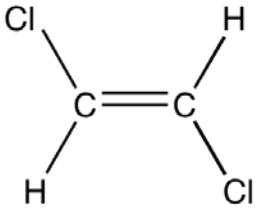
Isomer Example: $C_2H_2Cl_2$ (dichloroethene) Isomers



1,1−dichloroethene



cis−1,2−dichloroethene



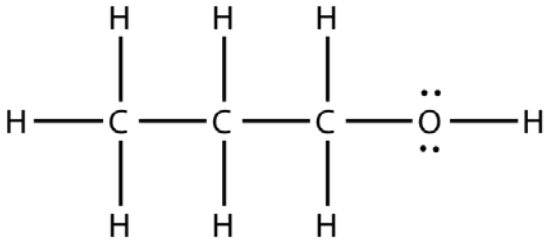
trans−1,2−dichloroethene

IMPORTANT: Resonance Structures are NOT isomers!

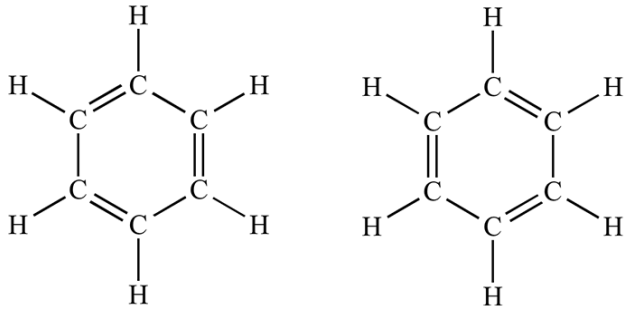
- The actual structure of a molecule with resonance is _____ hybrid amalgamation of that molecule.
- Isomers are MULTIPLE, DIFFERENT structures that can have _____ properties.

Let's Practice!

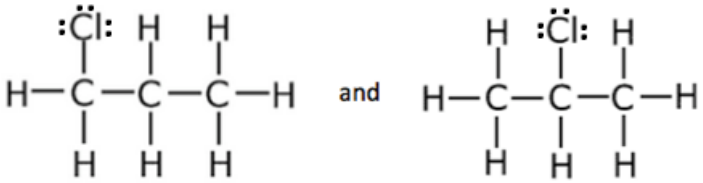
1. There are only 3 different isomers of C_3H_8O. One of them is shown below: can you draw the other two?



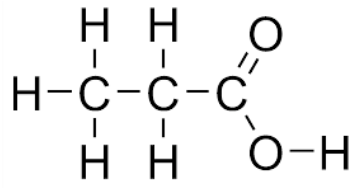
2. Are the two structures shown below isomers of each other or resonance structures? Justify your answer.



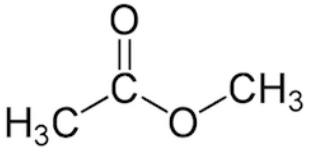
3. Are the two structures shown below isomers of each other or resonance structures? Justify your answer.



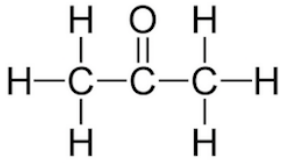
4. Which of the following structures is a structural isomer to the molecule shown below?



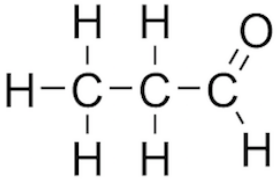
a.



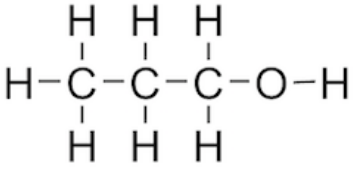
c.



b.



d.



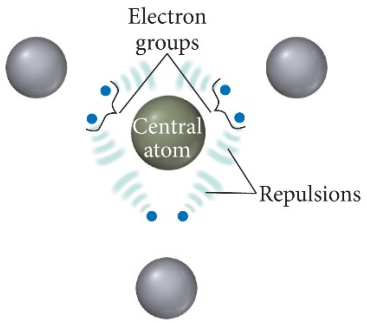
Challenge: Resonance Practice! Draw at least two valid resonance structures for each species below. Label formal charges on each structure.

CO_3^{2-}
HNO_3
O_3

VSEPR Theory: Valence Shell Electron Pair Repulsion

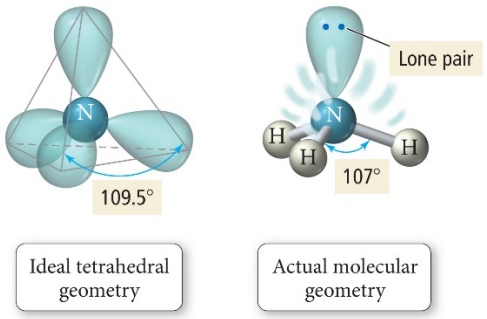
VSEPR Theory: (Valence Shell Electron Pair Repulsion) predicts the _____(shape) of molecules.

- Electrons surrounding an atom tend to _____ (repel) each other.



- Electrons (bonds and lone pairs) surrounding an atom will adopt a shape to _____ this repulsion, by arranging themselves as _____ as possible from each other.
- _____ and _____ bonds are treated the same as a SINGLE bond in terms of ability to repel electrons.

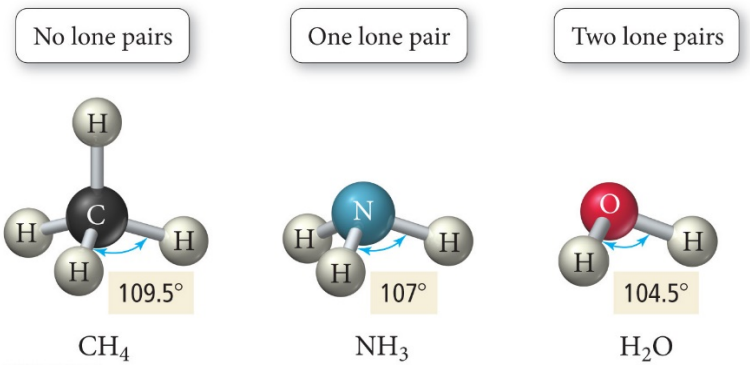
Magnitude of repulsion: **Lone pair > bonding pair(s)**



But why?

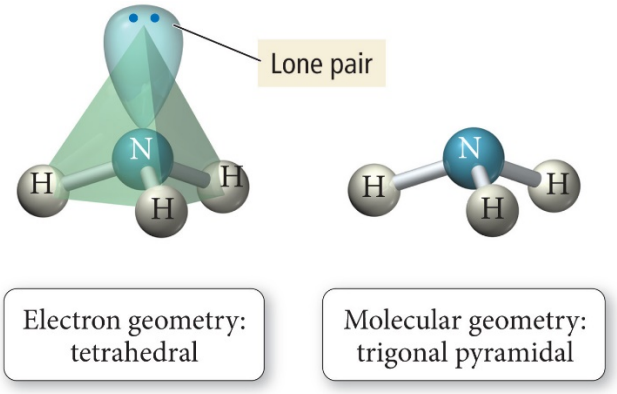
- Since lone pairs experience an attraction (or pull) from only ____ nucleus (as opposed to ____ nuclei for bonding electron pairs), lone pairs have a more concentrated electron density and thus a GREATER repulsive effect: they take up _____ space around an atom!

Effect of Lone Pairs on Molecular Geometry



Key Definitions

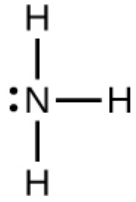
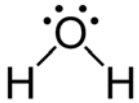
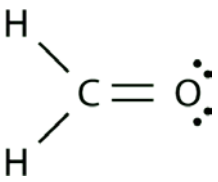
Electron Domains	Bonding Domains
Regions of electron _____ about a central atom.	Regions of _____ electron density about a central atom.
Used to determine <u>electron geometry</u>: • bond angles!	Used to determine <u>molecular geometry</u>: • VSEPR shape!
One electron domain equals: • One lone pair, or • One single bond, or • One double bond, or • One triple bond Yep, you read that right → a single, double, or triple bond only counts as **ONE** electron domain!	**One bonding domain** equals: • One single bond, or • One double bond, or • One triple bond Yep, you read that right → a single, double, or triple bond only counts as **ONE** bonding domain!



Important things to know about Molecular and Electron Geometry:

1. When lone pairs are _____ present on the central atom, molecular and electron geometries are the _____!
2. In expanded octets, lone pairs can be in two possible locations: <u>axial</u> (top and bottom) or <u>equatorial</u> (around the center). **Place lone pairs wherever they'll be _____ apart (maximum possible angle)!**

Let's Practice! How many electron domains and bonding domains are around the central atom for each of the following molecules?

1)	2)	3)
		
# electron domains? _____ # bonding domains? _____	# electron domains? _____ # bonding domains? _____	# electron domains? _____ # bonding domains? _____

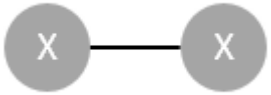
A Summary of Electron and Molecular Geometries

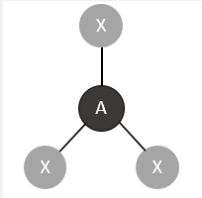
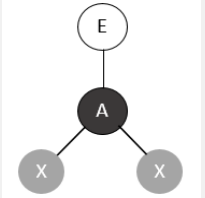
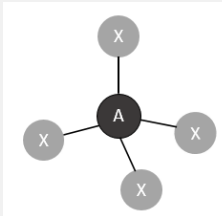
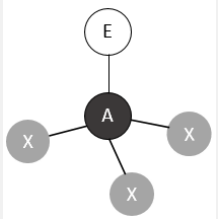
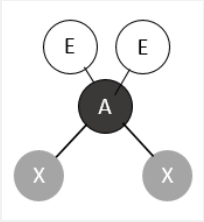
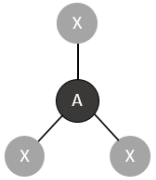
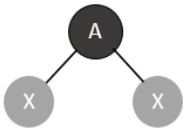
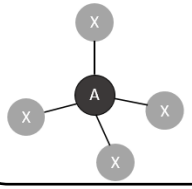
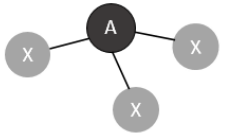
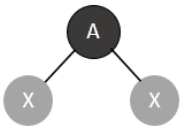
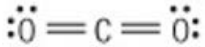
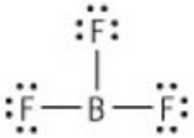
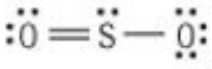
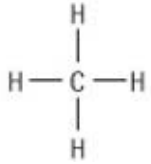
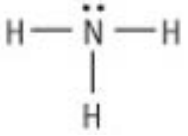
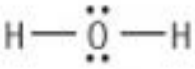
<u>You MUST be able to identify VSEPR shapes</u> of Lewis structures by their correct names!

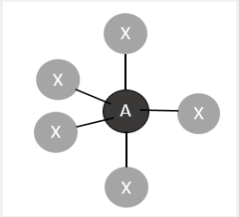
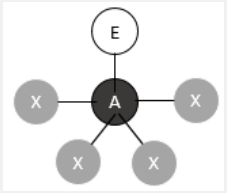
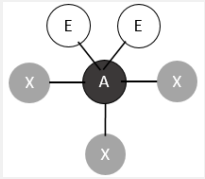
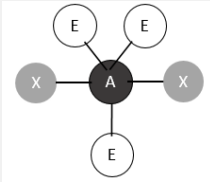
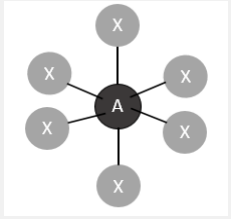
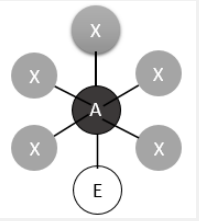
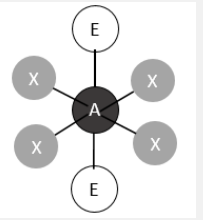
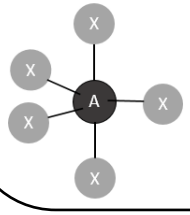
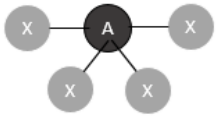
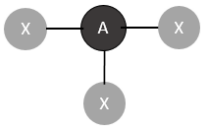
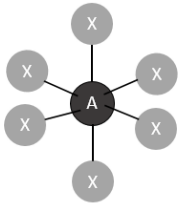
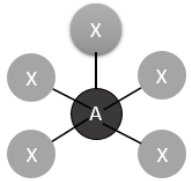
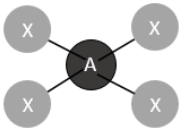
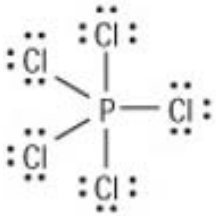
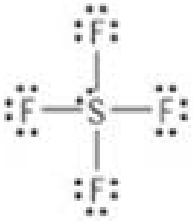
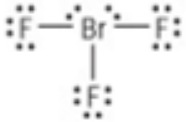
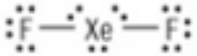
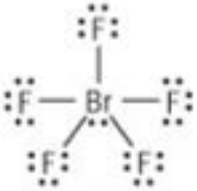
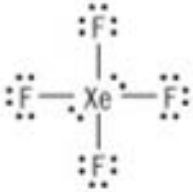
- Count electron groups around the central atom
- Each of the following is considered one electron group
  - A lone pair, a single bond, a double bond, a triple bond
- Electron groups repel each other to get as far apart as possible

A = central atom, **X** = terminal atom, **E** = a lone pair of electrons

Only 2 atoms = Linear

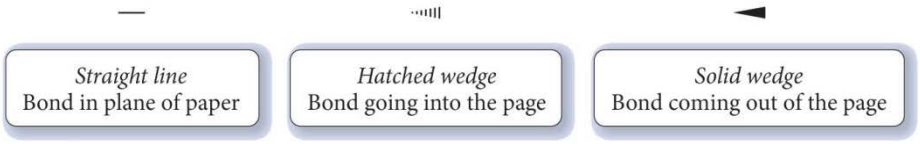


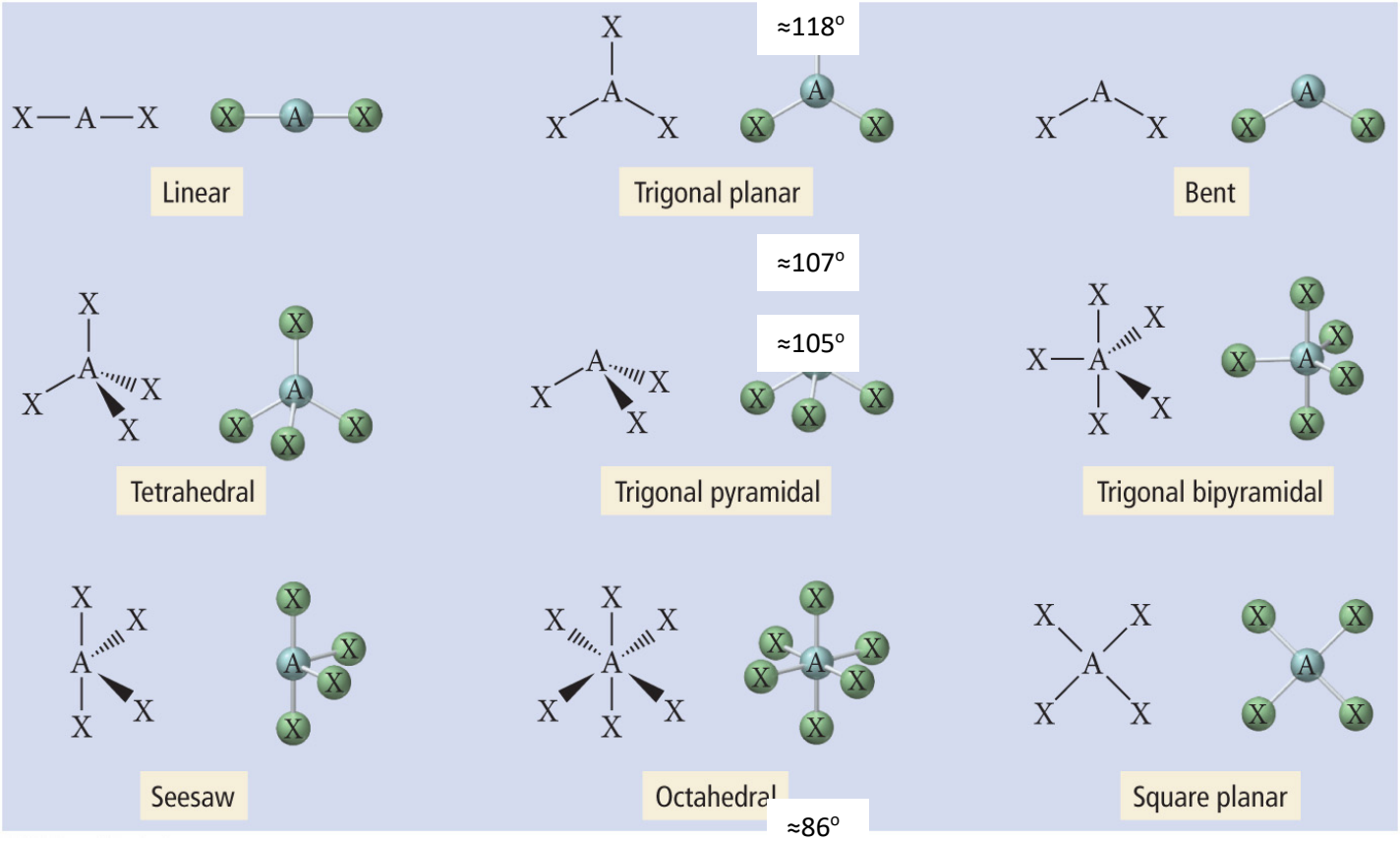
# of e⁻ groups	2	3		4		
General Formula	AX_2	AX_3	AX_2E	AX_4	AX_3E	AX_2E_2
Electron Geometry	Linear 	Trigonal planar 	Trigonal planar 	Tetrahedral 	Tetrahedral 	Tetrahedral 
Molecular Geometry (VSEPR Shape)	Linear 	Trigonal planar 	Bent 	Tetrahedral 	Trigonal pyramidal 	Bent 
Bond Angle	180°	120°	< 120°	109.5°	< 109.5°	<< 109.5°
		More e- repulsion = smaller angle		More e- repulsion = smaller angle		
Example						

# of e⁻ groups	5				6		
General Formula	AX_5	AX_4E	AX_3E_2	AX_2E_3	AX_6	AX_5E	AX_4E_2
Electron Geometry	Trigonal bipyramidal 	Trigonal bipyramidal 	Trigonal bipyramidal 	Trigonal bipyramidal 	Octahedral 	Octahedral 	Octahedral 
Molecular Geometry (VSEPR Shape)	Trigonal bipyramidal 	See-saw 	T-shaped 	Linear 	Octahedral 	Square pyramidal 	Square planar 
Bond Angle	120 ° (equatorial) 90° (axial)	< 120° equatorial) < 90° (axial)	< 90°	180°	90°	< 90°	90°
Example							

How to Draw 3-D Molecule Shapes on Paper

You <u>won't</u> be expected to use this 3-D drawing strategy, but <u>you have to know what it means when you see it</u> with an example molecule!





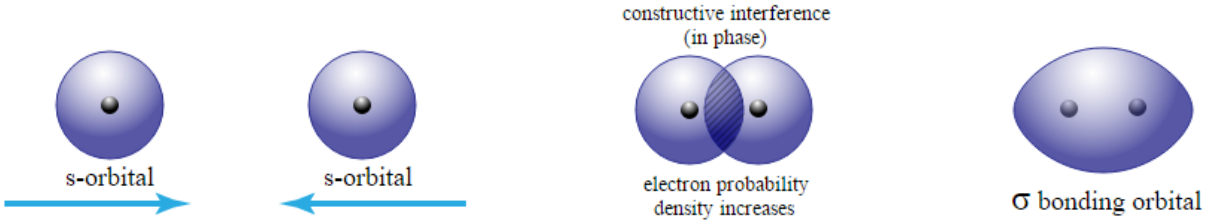
Let's Practice! Complete the table below.

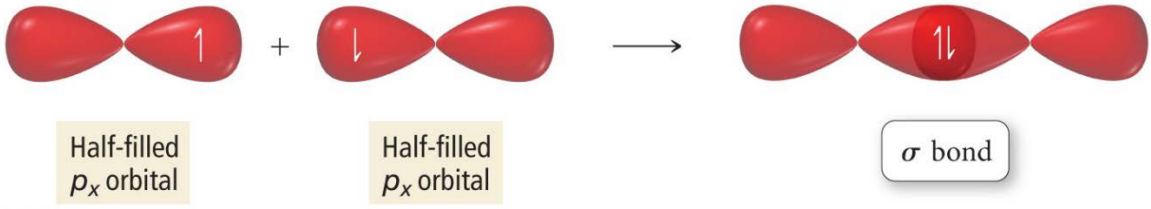
	Lewis Dot Structure (2D) (w/ predicted bond angles)	**Electron geometry**	**Molecular geometry**
CH_2O		# electron domains: ___ e^- geometry:	# bonding domains: ____ VSEPR shape:

H_3O^+		# electron domains: ___ e^- geometry:	# bonding domains: ____ VSEPR shape:
ClF_3		# electron domains: ___ e^- geometry:	# bonding domains: ____ VSEPR shape:
CO_2		# electron domains: ___ e^- geometry:	# bonding domains: ____ VSEPR shape:
XeF_4		# electron domains: ___ e^- geometry:	# bonding domains: ____ VSEPR shape:
IF_5		# electron domains: ___ e^- geometry:	# bonding domains: ____ VSEPR shape:
PF_5		# electron domains: ___ e^- geometry:	# bonding domains: ____ VSEPR shape:
$TeCl_4$		# electron domains: ___ e^- geometry:	# bonding domains: ____ VSEPR shape:

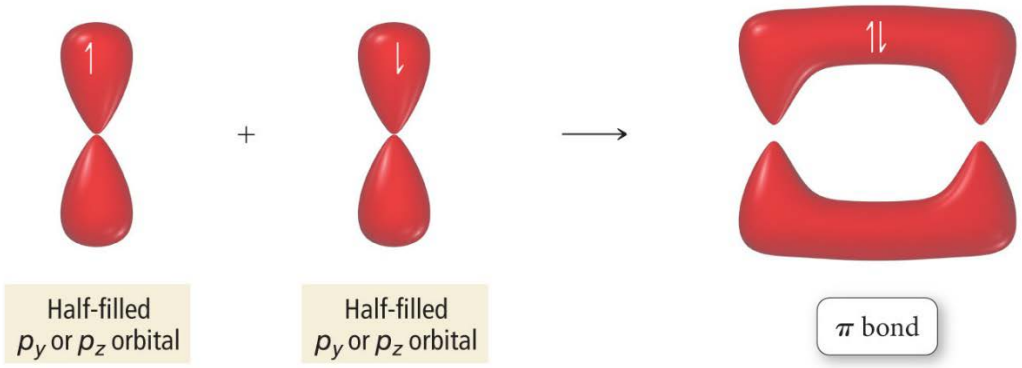
Sigma (σ) and Pi (π) Bonding

Sigma bond: covalent bond formed by orbitals overlapping end to end. The electron density is concentrated between the nuclei of the two atoms involved in a bond.



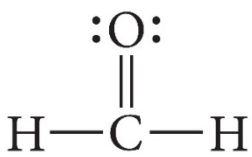


Pi bond: covalent bond formed by orbitals overlapping side by side. The electron density is concentrated above and below the nuclei of the two atoms involved in a bond.

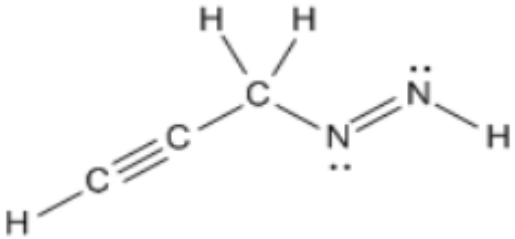


Bond Type	Made of	Length / Strength
single bond	____ sigma bond	longest/ weakest
double bond	____ sigma bond + ____ pi bond	medium length/ strength
triple bond	____ sigma bond + ____ pi bonds	shortest/ strongest

Let's Practice! Identify the number of sigma and pi bonds in each structure shown below:



sigma bonds _____
pi bonds _____



sigma bonds _____
pi bonds _____

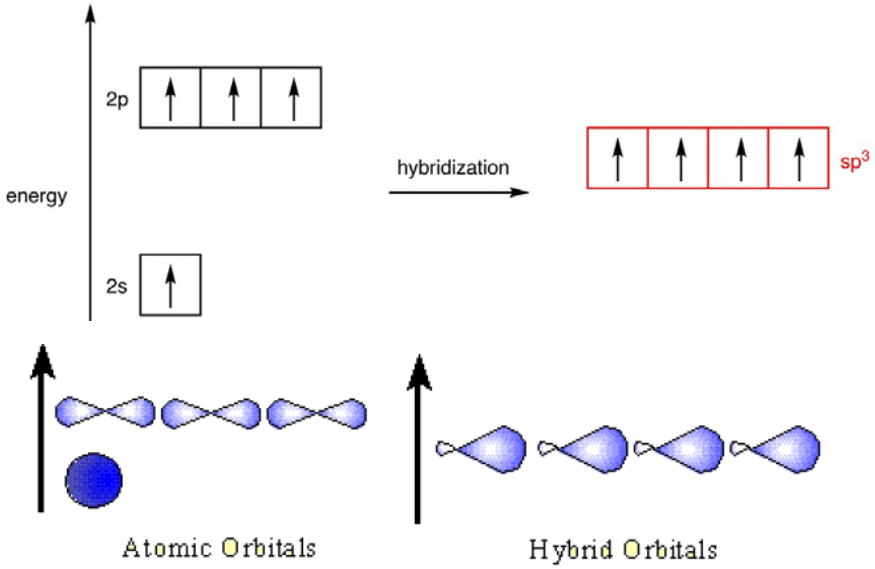
Hybridization

Hybrid Orbitals: orbitals of equal energy created by _____ two or more valence orbitals on the same atom

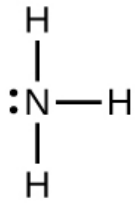
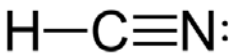
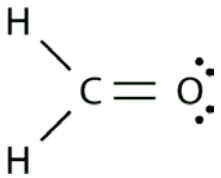
- The same type of atom can have _____ types of hybridization depending on the atoms it is bonded with.
- Hybridization can be determined by counting regions of _____ density: electron domains!

# of Electron Domains	# of Hybrid Orbitals	Electron Geometry	Hybridization
2	2	Linear	*sp*
3	3	Trigonal planar	sp^2
4	4	Tetrahedral	sp^3
5	5	Trigonal bipyramidal	sp^3d
6	6	Octahedral	sp^3d^2

Example: Carbon in CH_4



Let's Practice! Identify the hybridization of the valence orbitals around the central atom in each molecule below.

1)  # electron domains? _____ hybridization? _____	2)  # electron domains? _____ hybridization? _____	3)  # electron domains? _____ hybridization? _____

Let's Practice!

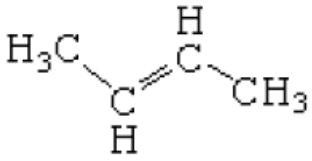
1. What hybridization change does the carbon atom undergo in the combustion of methane?

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$$

   a. $sp \rightarrow sp^2$
   b. $sp^2 \rightarrow sp^3$
   c. $sp^3 \rightarrow sp$
   d. $sp^2 \rightarrow sp$

2. How many sigma and pi bonds are present in the following molecule?

   a. 8 sigma bonds and 1 pi bond
   b. 8 sigma bonds and 2 pi bonds
   c. 10 sigma bonds and 2 pi bonds
   d. 11 sigma bonds and 1 pi bond

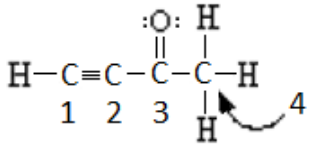


CCl_4, CO_2, PCl_3, PCl_5, SF_6

3. Which of the following does not describe any of the molecules above?

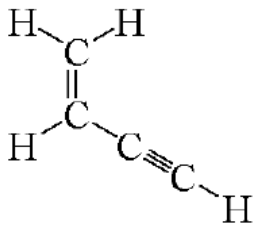
   a. tetrahedral
   b. linear
   c. octahedral
   d. square planar

4. Label the hybridization at each carbon in the molecule below.



	C1	C2	C3	C4
a.	sp	sp	sp^3	sp^3d
b.	sp	sp	sp^2	sp^3
c.	sp	sp^2	sp^2	sp^2
d.	sp^2	sp^2	sp^3	sp^3

5. How many sigma bonds and pi bonds are in the following molecule?



a. five σ and two π
b. five σ and three π
c. five σ and five π
d. seven σ and two π
e. seven σ and three π

6. Complete the chart below:

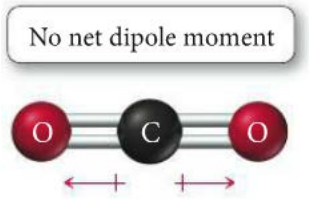
$HCCCH_3$
Lewis structure: **Label** the hybridization AND molecular geometry around each carbon in the molecule above. Total # of sigma bonds: _____ Total # of pi bonds: _____
$H_2CCH - O - CH_3$
Lewis structure: **Label** the hybridization AND molecular geometry around each carbon and oxygen in the molecule above. Total # of sigma bonds: _____ Total # of pi bonds: _____

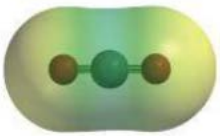
Molecular Polarity

Just like bonds, molecules can be polar or non-polar.

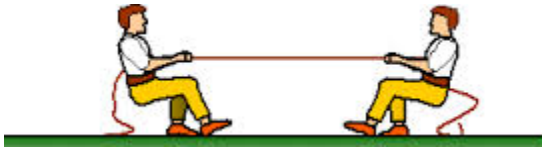
Non-polar molecules: (**dipole moment =** _____) either the bonds are non-polar, or the bond dipoles cancel out!

Examples: CO_2, CH_4, SiF_4



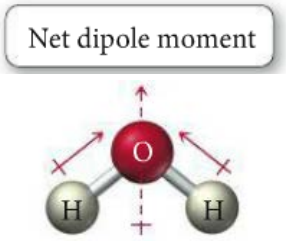


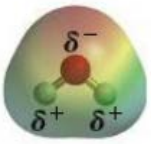
Think of dipoles like _____: if two forces are pulling on the same object in **equal** but **opposite** directions, the object does _____ move → non-polar molecule!



Polar Molecules: Unequal distribution of electron density, because bond dipoles don't cancel , the presence of electron pairs, or both.

Examples: HCl, H_2O, CH_2F_2, PH_3





"Spatial asymmetry" of electron density = polar!

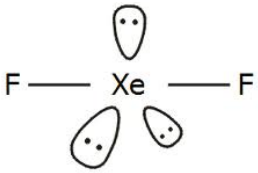
Determining Molecular Polarity

1. **If** _____ _____ **are present on the central atom**, the molecular is **typically** _____: lone pairs on the central atom make the molecule polar because their presence creates increased electron repulsion and thus, an unequal distribution of electron density.

   However, there are the following **exceptions**:

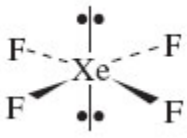
   a. **Trigonal bypyramidal** structures with three lone pairs (linear VSEPR).

      Example: XeF_2



   b. **Octahedral structures** with two lone pairs (square planar VSEPR)

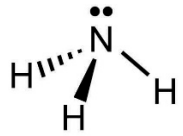
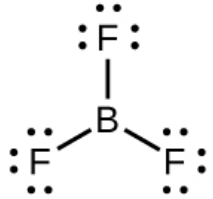
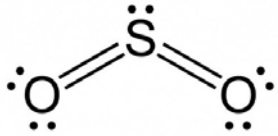
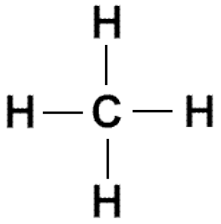
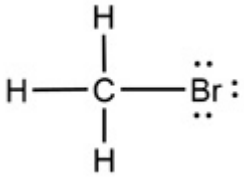
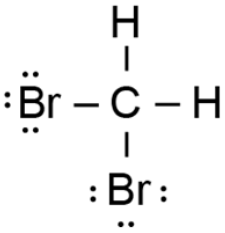
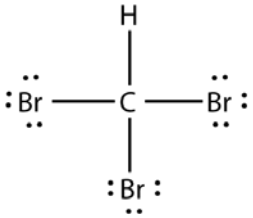
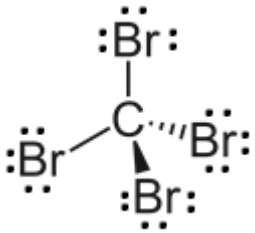
      Example: XeF_4



2. If lone pairs are _____ present on the central atom, and:
   a. **terminal** (non-central) **atoms are all the** _____**,** the molecule is _____.
   b. **terminal** (non-central) **atoms are different,** the molecule is _____.

Let's Practice!

Given the Lewis dot structures below, determine if the molecule is polar or non-polar.

1)  Is the molecule: polar or non-polar?	2)  Is the molecule: polar or non-polar?	3)  Is the molecule: polar or non-polar?	4)  Is the molecule: polar or non-polar?
5)  Is the molecule: polar or non-polar?	6)  Is the molecule: polar or non-polar?	7)  Is the molecule: polar or non-polar?	8)  Is the molecule: polar or non-polar?

And… Even More Practice!

For each of the compounds given below, identify its VSEPR shape, bond angle(s), the hybridization of its central atom, and its molecular polarity.

CO	**NH_4^+**	**H_2S**
Lewis structure:	Lewis structure:	Lewis structure:
Molecular geometry:_____ Bond angle(s):_____ Central Atom Hybridization_____ Polar or not?_____	Molecular geometry:_____ Bond angle(s):_____ Central Atom Hybridization_____ Polar or not?_____	Molecular geometry:_____ Bond angle(s):_____ Central Atom Hybridization_____ Polar or not?_____

PF_5	**IF_4^-**	**SeF_4**
Lewis structure:	Lewis structure:	Lewis structure:
Molecular geometry:_____ Bond angle(s):_____ Central Atom Hybridization_____ Polar or not?_____	Molecular geometry:_____ Bond angle(s):_____ Central Atom Hybridization_____ Polar or not?_____	Molecular geometry:_____ Bond angle(s):_____ Central Atom Hybridization_____ Polar or not?_____
SiF_4	**CO_2**	**XeF_3^+**
Lewis structure:	Lewis structure:	Lewis structure:
Molecular geometry:_____ Bond angle(s):_____ Central Atom Hybridization_____ Polar or not?_____	Molecular geometry:_____ Bond angle(s):_____ Central Atom Hybridization_____ Polar or not?_____	Molecular geometry:_____ Bond angle(s):_____ Central Atom Hybridization_____ Polar or not?_____

Intro to IMFs: Sooooo Attractive!

Intramolecular forces: the attraction that results when electrons are given, taken, or shared to form _____ bond.

- Ionic bonds
- Covalent bonds
- Metallic bonds

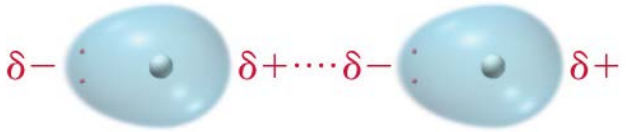
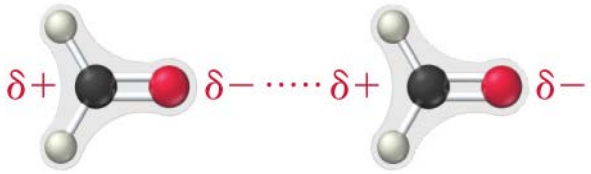
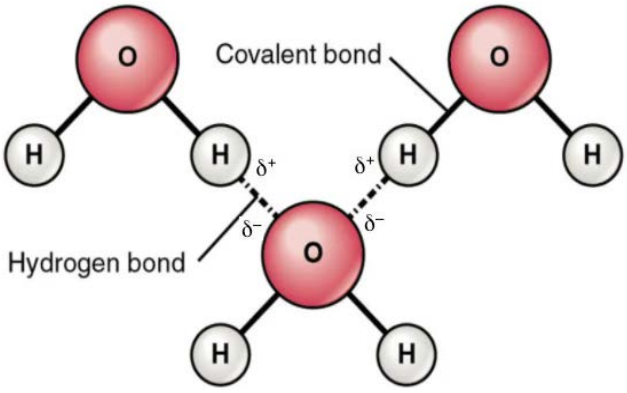
Intermolecular forces (IMFs): the attraction _____ two or more distinct molecules or particles.

- In a physical change, only _____ forces are broken because no chemical bonds are broken (same substance).
- In a chemical change, _____ forces are broken as chemical bonds are broken/formed (new substance).

Intramolecular forces are much, MUCH stronger than intermolecular forces!

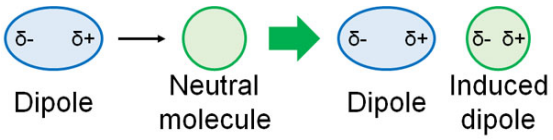
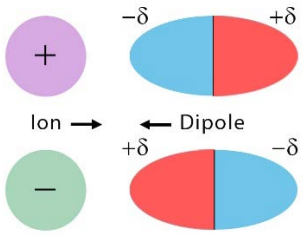
Think of the children!

Types of IMFs

Type	Present in	Molecular perspective
London Dispersion Forces/ Van der Waals Forces (LDFs)	• _____ molecules and atoms (nonpolar and polar) • Strength depends on total _____ of electrons in an atom/molecule	 instantaneous dipoles
Dipole-dipole attractive forces	• Between _____ molecules (which already have a dipole moment)	
Hydrogen "bonding" attractive forces	• Between H bonded to F, O, or N on one molecule and **lone pair** on F, O, or N of another molecule → so much _____ ! →Hydrogen "bonds" are _____ actual bonds (intramolecular forces); the name is very misleading!	

If comparing **similar-sized** particles: **Hydrogen bonding IMFs > dipole-dipole IMFs > LDFs**

Mixed IMFs! Attractive Forces Between Two _____ Compounds

Type	Present in	Molecular perspective
Dipole – induced dipole attractive forces	**Polar** molecules combined with **Nonpolar** molecules	
Ion – dipole attractive forces	**Ions/ Ionic** Compounds combined with Polar molecules	

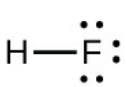
If comparing **similar-sized** particles:

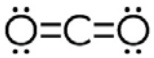
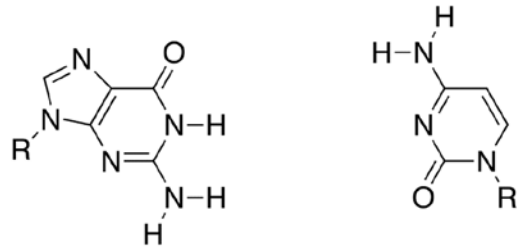
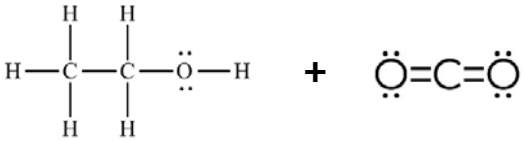
Ion – Dipole **>** Hydrogen bonding **>** Dipole – dipole **>** Dipole – Induced Dipole **>** LDFs

In Summary

	Non-polar, single elements	*polar + non-polar*	polar no H – FON	polar with H – FON	*polar + ions*
London dispersion forces (LDFs or Van der Waals)					
Dipole – Induced Dipole (multiple compounds)					
Dipole – Dipole					
Hydrogen "bonding" attractive forces					
Ion – Dipole (multiple compounds)					

Let's Practice!

Molecule 1. Label each as polar, non-polar, or an ion 2. If polar, is it FON?	**IMFs present?** 1. Which IMFs are present between the compounds? 2. Underline the dominant (strongest) IMF exhibited
$CH_3CH_2CH_3$	
	

	
	
 Guanine + Cytosine	
	

Multiple Choice Practice

1. The forces of attraction between the permanent partially positive regions of one molecule and permanent partially negative regions of another nearby molecule are called _____.

   a. covalent bonds
   b. hydrogen bonding
   c. London dispersion forces
   d. dipole−dipole forces

2. A polar molecule is one in which:

   a. ions exist
   b. only london dispersion forces exist
   c. it has a VSEPR structure that is symmetrical
   d. electron density is unequally distributed

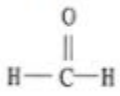
3. Which of the following processes involves breaking intermolecular forces?

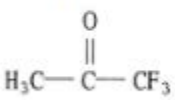
   a. $H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$
   b. $2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 4\ CO_2(g) + 6\ H_2O(g)$
   c. $I_2(g) \rightarrow 2\ I(g)$
   d. $H_2O(s) \rightarrow H_2O(l)$

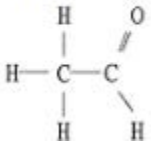
4. What is the dominant intermolecular force in CH_3OH?

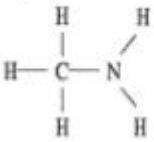
   a. London dispersion forces
   b. Ion-dipole attraction
   d. Dipole-dipole attraction
   e. Hydrogen bonding

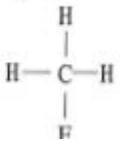
5. Which one of the following substances will have hydrogen bonding as one of its intermolecular forces?

(A) 

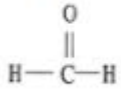
(B) 

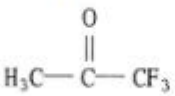
(C) 

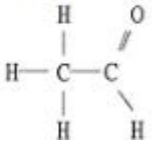
(D) 

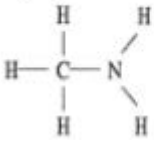
(E) 

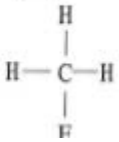
6. Circle ALL of the following substances which <u>cannot</u> hydrogen bond with another molecule of the **same** substance, but would be capable of hydrogen bonding with a **different** molecule that has H directly bonded to fluorine, oxygen, or hydrogen.

(A) 

(B) 

(C) 

(D) 

(E) 

The Language of IMFs (i.e. How to FR)

Notes about Language: Talking about IMFs can be tricksy! And of course, IMF-based questions are VERY common on free response portion of the AP test, so it's essential to carefully choose your language to avoid losing points.

Most IMF points are lost when student inadvertently imply (or, worse, directly state) that during state changes:

1. _____ bonds are being broken during state changes **(NOPE!)**
2. Molecules are being _____ apart **(NOPE!)**

How to Talk about IMFS	How to Talk about Bonds
• _____ IMFs o Avoid breaking IMFs • _____ molecules o NEVER break apart molecules • IMFs are _____ molecules	• _____ bonds • Bonds are _____molecules

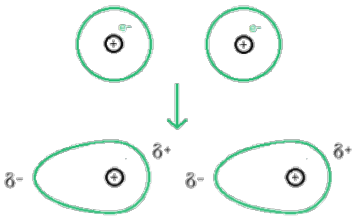
Free Response Practice: What **<u>type(s)</u>** of intermolecular OR intramolecular attractive forces must be overcome for each of the following processes to occur?

1. $CO_2(g) \rightarrow C(s) + O_2(g)$, $\Delta H = 393.5$ kJ/mol

2. $CO_2(s) \rightarrow CO_2(g)$, $\Delta H = 25.2$ kJ/mol

A Closer Look at London Dispersion Forces (LDFs): Induced Dipole – Induced Dipole Attraction

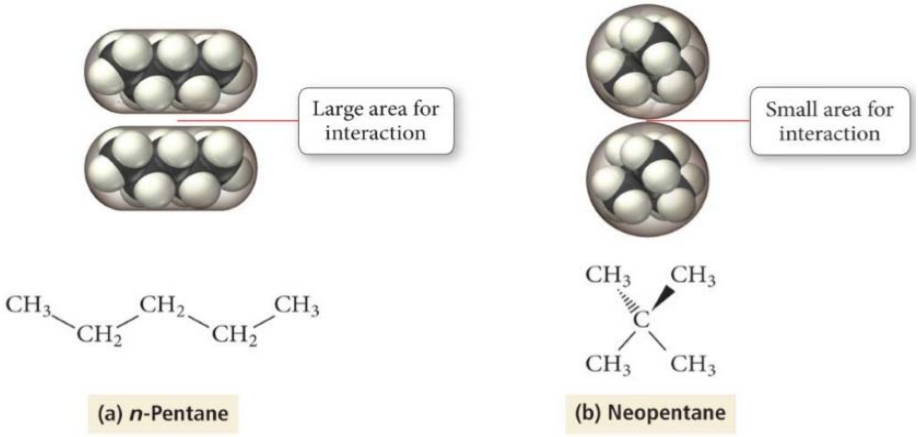
LDFs are determined by the _____ of a molecule (i.e. how much the electron cloud can temporarily be shifted)



- ➔ an electron cloud, even in a nonpolar molecule, can _____ shift, causing one side of the molecule to be more negative than another
- ➔ this temporary dipole can _____ (cause) a temporary dipole on a neighboring molecule (hence the name "induced dipole")

What can increase LDFs?

1. Greater electron cloud (more electrons): molecule is _____ polarizable = ____ LDFs
2. Increase in molar mass (implies more electrons): molecule is _____ polarizable = ____ LDFs
3. Increase in surface-to-surface contact area: ____ induced dipole = ____ LDFs



Boiling point: 36.1°C Boiling point: 10.0°C

Be careful!

- When non-polar substances with _____ London dispersion forces have a considerably _____ (and thus very polarizable) electron cloud than the polar molecules, the LDFs can be quite substantial and be **STRONGER** than hydrogen bonding forces or dipole-dipole forces (!!)

Example: Cl_2 has a higher boiling point that HCl. Explain.

Let's Practice!

1. Rank the following in order of increasing LDFs: CH_3CH_3, CH_4, $CH_3CH_2CH_3$

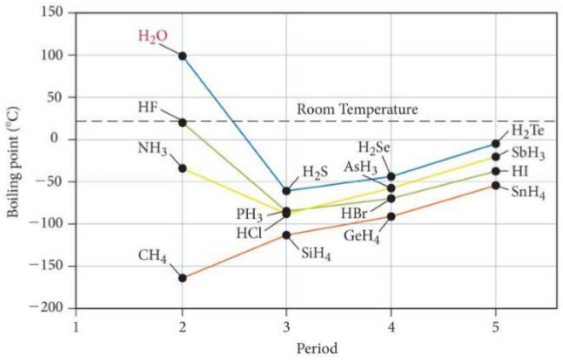
2. Rank the following in order of increasing LDFs: Br_2, F_2, Cl_2, I_2

A Closer Look at Hydrogen Bonding Attractive Forces

Note: Hydrogen "bonds" are _____actual bonds (intramolecular forces), and thus the name is very misleading!

Hydrogen "Bonding": force of attraction between hydrogen atom bonded to a small highly electronegative atom (F, O, and N) and the unshared electron pair on another electronegative atom of F, O, or N.

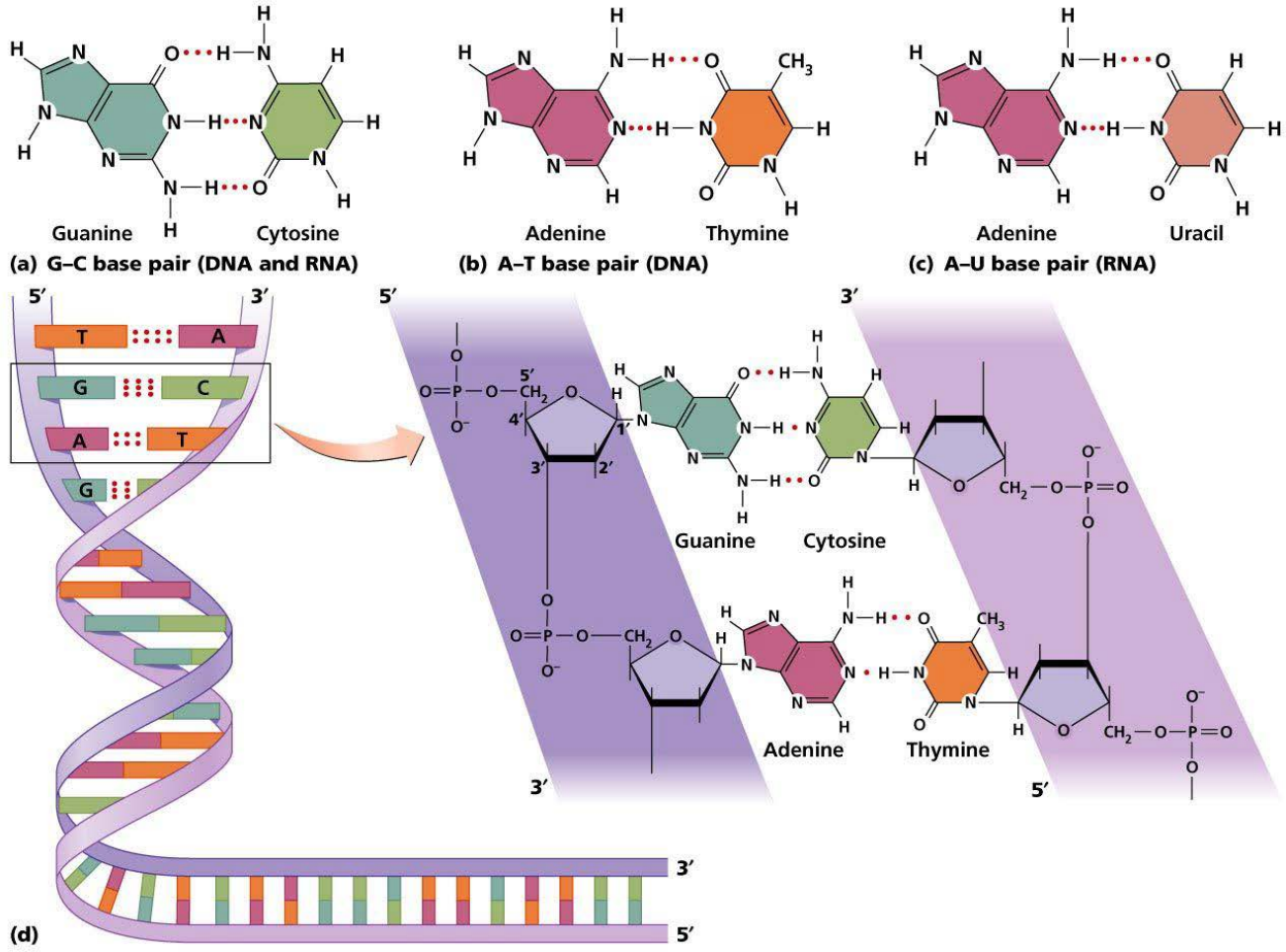
- Hydrogen "bonding" is F-O-N! ☺
- Hydrogen "bonding" is usually depicted with a dotted or dashed line.
- Hydrogen "bonding" is responsible for some of the unique properties of water, including its relatively _____ boiling point.



How do you represent Hydrogen Bonding? Let's look at NH_3.

Hydrogen Bonding: It's in Your DNA!

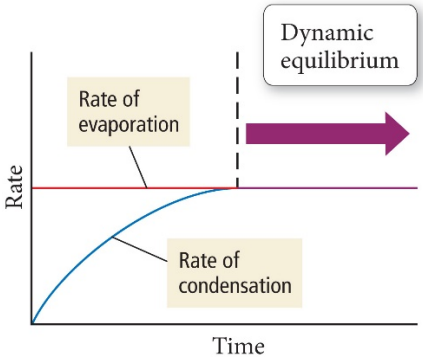
The different number of hydrogen bonds in each complementary base pair (adenine and thymine vs cytosine and guanine) helps ensure that the base pairs will match up correctly!

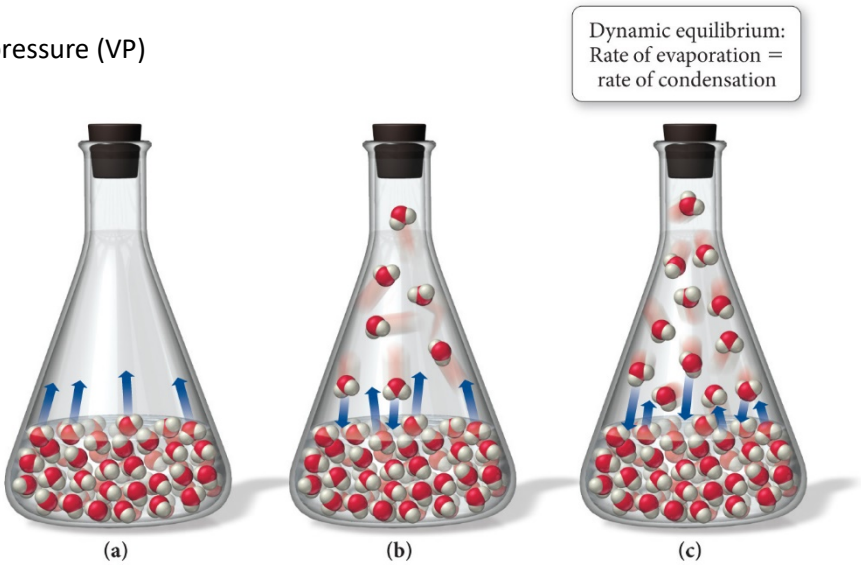


IMFs in Action

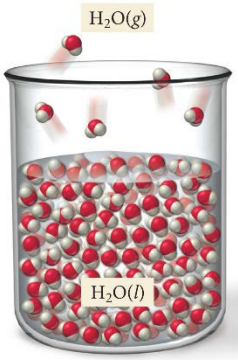
Measure of Intermolecular Forces

1. Vapor Pressure: the pressure exerted by a gas (vapor) when it is in dynamic equilibrium with its liquid (must be in _____ container!)
    a. The weaker the attractive forces between the molecules, the more molecules will be in vapor (and vice versa).
    b. Thus, ____ IMFs = ____ vapor pressure (VP)





2. Volatility: how quickly a substance evaporates.
    a. The weaker the attractive forces between the molecules, the more quickly and easily molecules will separate from each other and enter the gas phase.
    b. Thus, ____ IMFs = ____ volatility



3. Boiling point: the temperature at which molecules separate from each other in the liquid phase and enter the gas phase.

4. Melting point: temperature at which molecules of a solid have enough thermal energy to overcome IMFs and become a liquid
    a. Thus, ____ IMFs = ____ melting point (MP) = ____ boiling point (BP)

5. Solubility in water: amount of a given substance that will _____ in water
    a. Strong interactions form between polar/ionic solute particles and polar solvent molecules as they mix → energetically favorable!
    b. Thus, ____ IMFs = ____ solubility in water

Other Properties You Must be able to Explain with IMFs: The **Higher** these are, the **Stronger** the IMFs!

1. Surface tension: energy required to increase the surface area of a liquid
2. Capillary action: spontaneous rising of a liquid in a narrow tube
3. Viscosity: resistance to flow

***Note: Only Vapor Pressure and Volatility have an Inverse Relationship with IMF strength!**

Let's Practice! Draw TWO Lewis dot structures for each of the following molecules. Then, draw dipoles where appropriate and determine the type(s) of intermolecular forces that exist between the molecules.

NH_4^+ Type(s) of IMFs?	CH_3Br Type(s) of IMFs?
H_2O Type(s) of IMFs?	CO_2 Type(s) of IMFs?

More Practice: Draw one EACH of the two compounds given in each example. Then, draw dipoles where appropriate and determine the type(s) of intermolecular forces that exist between the molecules.

NH_4^+ and H_2O Type(s) of IMFs?	CH_3OCH_3 and H_2O Type(s) of IMFs?

More Practice!

1. What **type** of intermolecular OR intramolecular attractive forces must be overcome for each process below?
   a. $O_2(l) \rightarrow O_2(g)$

   b. $Al_2O_3(s) \rightarrow 2\ Al(s) + 3/2\ O_2(g)$

2. What is the predominant intermolecular force in CBr_4?
   a. London dispersion forces
   b. Ionic bonding
   c. Dipole-dipole attraction
   d. Hydrogen bonding
   e. Ion-dipole attraction

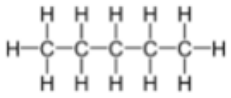
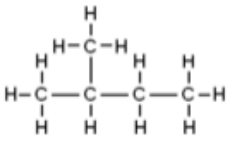
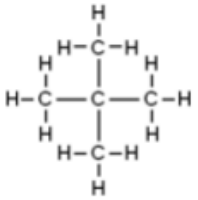
3. Which one of the following derivatives of methane has the highest boiling point?
   a. CBr_4 b. CCl_4 c. CF_4 d. CH_4

4. Based on the molecular mass and dipole moment of the five compounds in the table below, which should have the lowest vapor pressure?

Substance	Molecular Mass (amu)	Dipole Moment (D)
Propane, $CH_3CH_2CH_3$	44	0.1
Dimethylether, CH_3OCH_3	46	1.3
Methylchloride, CH_3Cl	50	1.9
Acetaldehyde, CH_3CHO	44	2.7
Acetonitrile, CH_3CN	41	3.9

   a. $CH_3CH_2CH_3$ b. CH_3OCH_3 c. CH_3Cl d. CH_3CHO e. CH_3CN

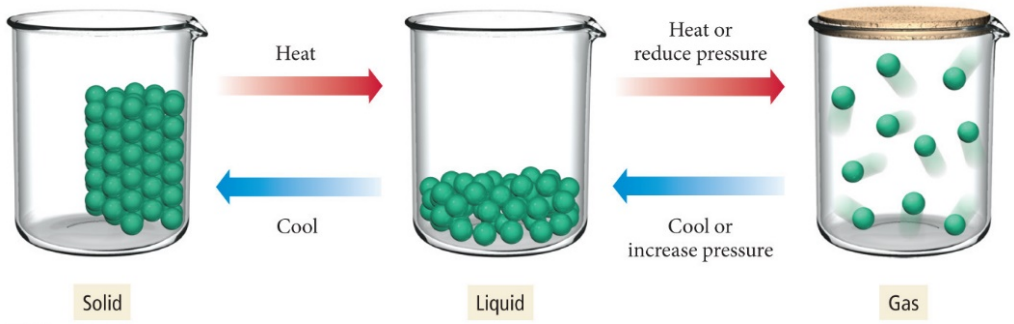
5. Which of the following BEST explains why neopentane has the lowest boiling point?

Common Name	*n*-pentane	isopentane	neopentane
Structure			
Formula	C_5H_{12}	C_5H_{12}	C_5H_{12}
Boiling Point °C	36.0	27.7	9.5

   a. Neopentane is less polarizable due to having fewer electrons.
   b. Neopentane is more polarizable due to having more electrons.
   c. Neopentane has the shortest carbon chains and thus the least surface area.
   d. Neopentane has the shortest carbon chains and thus the most surface area.

IMFs in Action = Heating Curves

KMT: The kinetic-molecular theory is based on the idea that particles of matter are always in motion.



KMT, IMFs and Changes of State: because attractive forces between the molecules are fixed, changing a material's state of matter require changing the amount of kinetic energy the particles have, or limiting their freedom.

1. **Gaseous state**: particles have _____ freedom of motion.
   a. Their kinetic energy _____ the attractive forces between the molecules.
2. **Liquid state**: particles have _____ freedom; they can move around a little within the liquid.
   a. They have enough kinetic energy to overcome _____ of the attractive forces, but not enough to _____ each other.
3. **Solid state**: particles are locked in place, they _____ move around.
   a. Although the particles _____, they do _____ have enough kinetic energy to overcome the attractive forces.

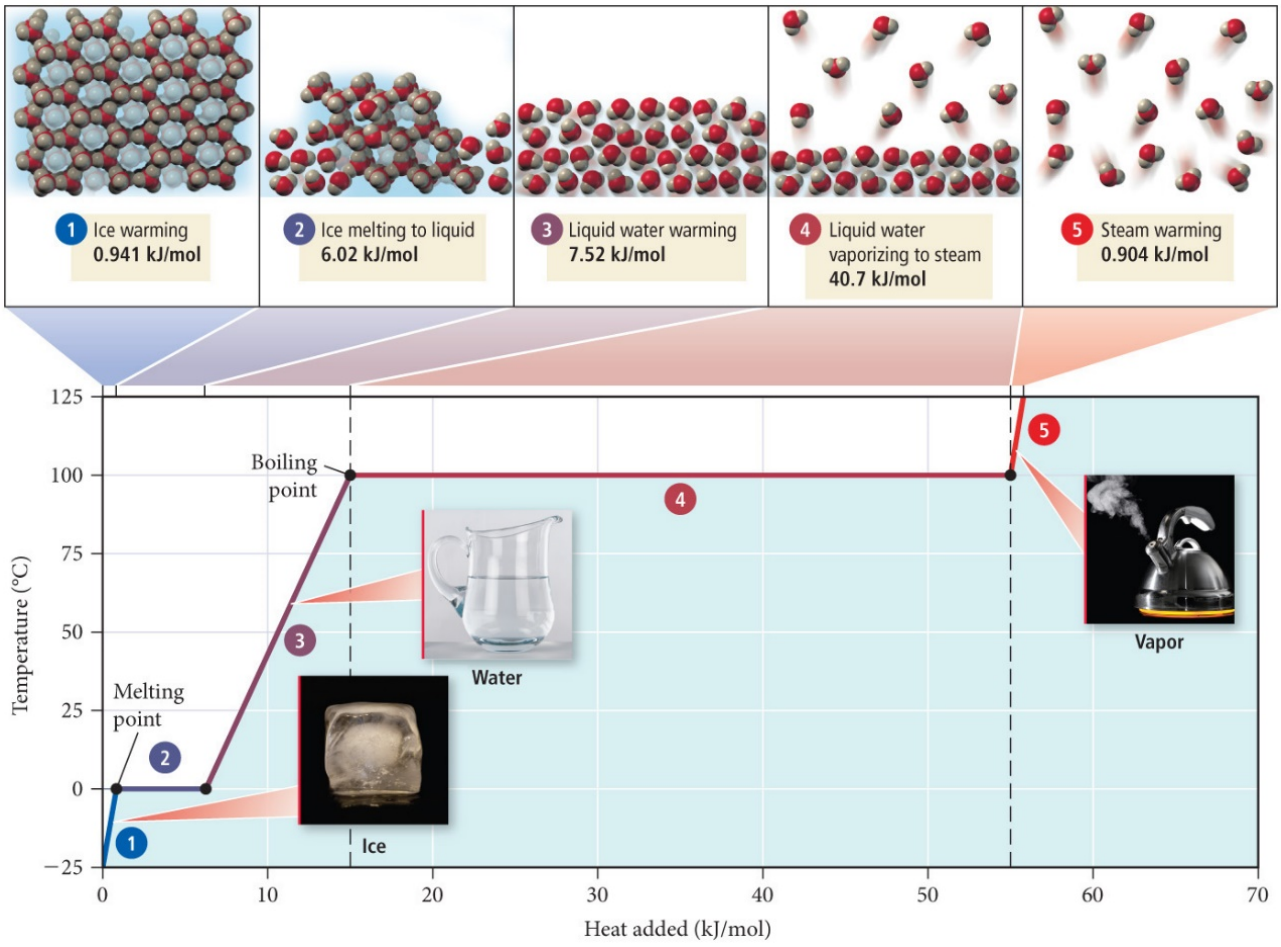
The strength of the attractive forces between particles of a substance determine its state!

- At room temperature, moderate to _____ attractive forces result in materials that are solids or liquids.
- The _____ the attractive forces, the _____ the boiling/ melting point!

State	Density	Shape	Volume	Strength of Intermolecular Forces (Relative to Thermal Energy)
Gas	Low	Indefinite	Indefinite	Weak
Liquid	High	Indefinite	Definite	Moderate
Solid	High	Definite	Definite	Strong

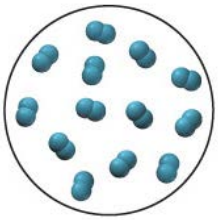
Heating and Cooling Curves

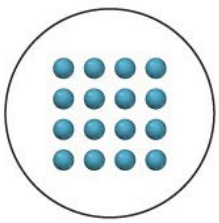
A graph of the temperature of the system versus the amount of heat added.

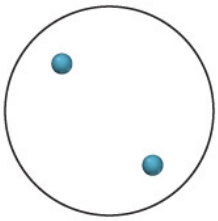


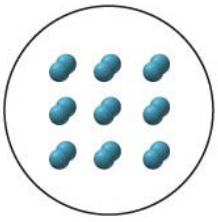
- In thermochemistry, our focus with heating and cooling curves is on how much _____ energy is required to change from one state of matter to another.
- <u>For this unit</u>, our focus is on two things:
  - the relative amount of _____ energy for each state of matter and
  - the strength of the _____ forces (IMFs) holding the particles together in that state

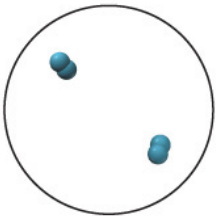
Let's Practice! If liquid nitrogen is shown in the image to the right, which of the images below best depicts nitrogen after it has boiled?



a) 

b) 

c) 

d) 

Solubility and IMFs

Remember solubility rules? ☺ The bolded 3 at the beginning are the only ones you need to memorize, but there are lots and lots of solubility patterns we can observe.

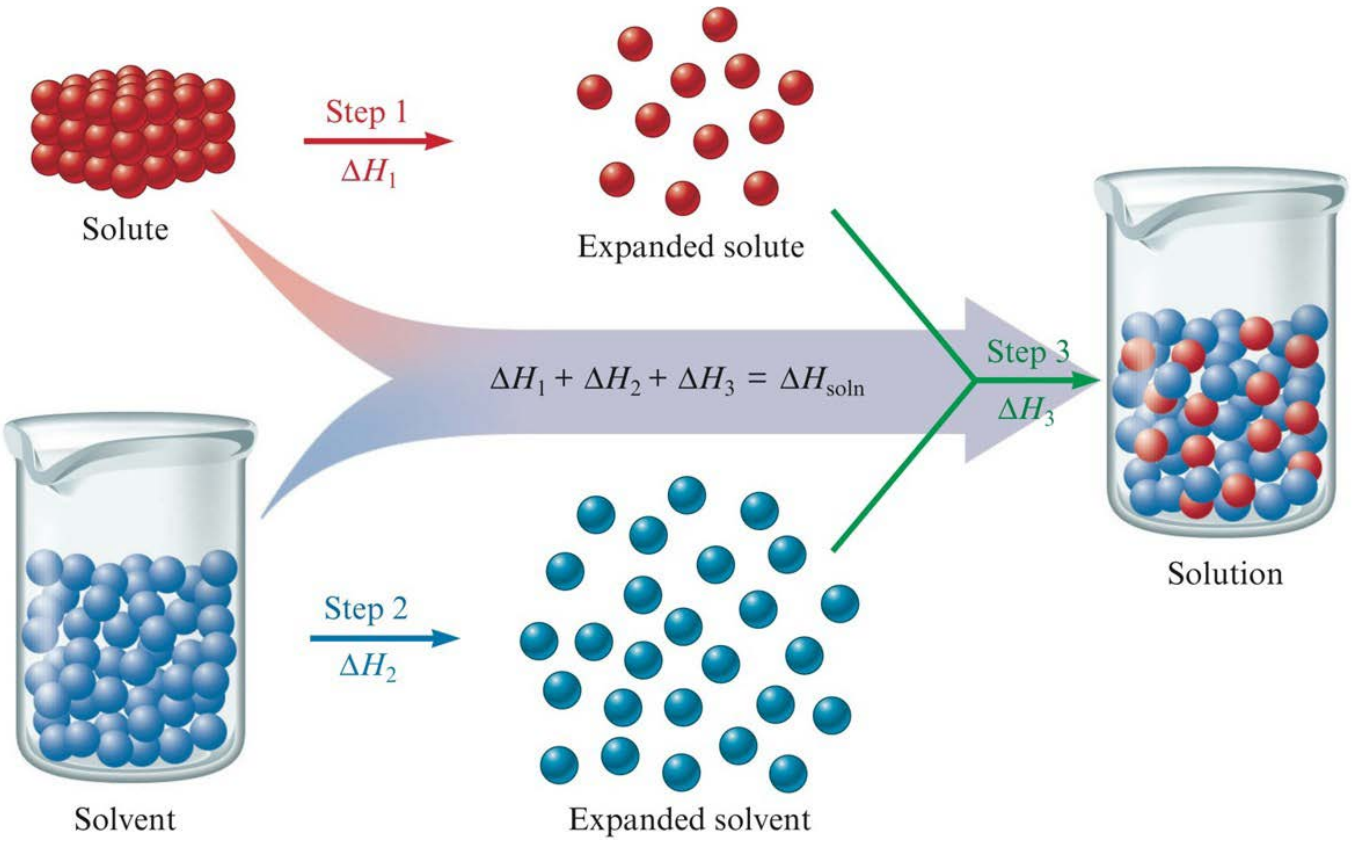
1. **Always soluble: _____ metal cations, NH_4^+, NO_3^-** (also ClO_3^-, ClO_4^-, $C_2H_3O_2^-$, HCO_3^-)
2. Generally soluble:
   a. Bromide, chloride, and iodide anions are soluble except when combined with Ag^+, Pb^{2+}, and Hg_2^{2+}.
   b. SO_4^{2-} is soluble except when combined with Sr^{2+}, Ba^{2+}, Pb^{2+}, and Hg_2^{2+}.
3. Generally insoluble:
   a. OH^- and S^{2-} are insoluble except when combined with Ca^{2+}, Sr^{2+}, Ba^{2+}, (and things from rule 1).
   b. CO_3^{2-}, PO_4^{3-}, SO_3^{2-}, and CrO_4^{2-} are insoluble except when combined with things from rule 1.

IMFs help explain these patterns of solubility!

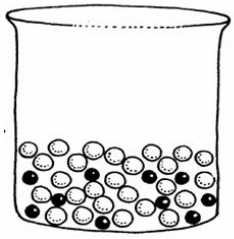
Dissolution depends on the forces of attraction between _____ and _____ particles

In order to dissolve a substance, you must:

1. Add energy: Overcome attractions (requires energy = endothermic) *"endo-ing" an attraction is endothermic!*
   a. Solute-solute IMFs (or ion-ion electrostatic attraction, if ionic)
   b. Solvent-solvent IMFs
2. Release energy: Form solute-solvent attractive forces upon mixing (releases energy = exothermic)

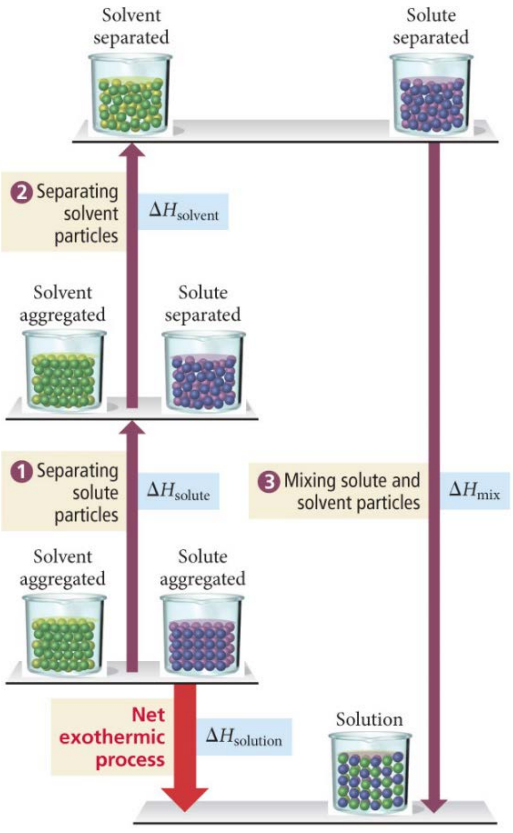
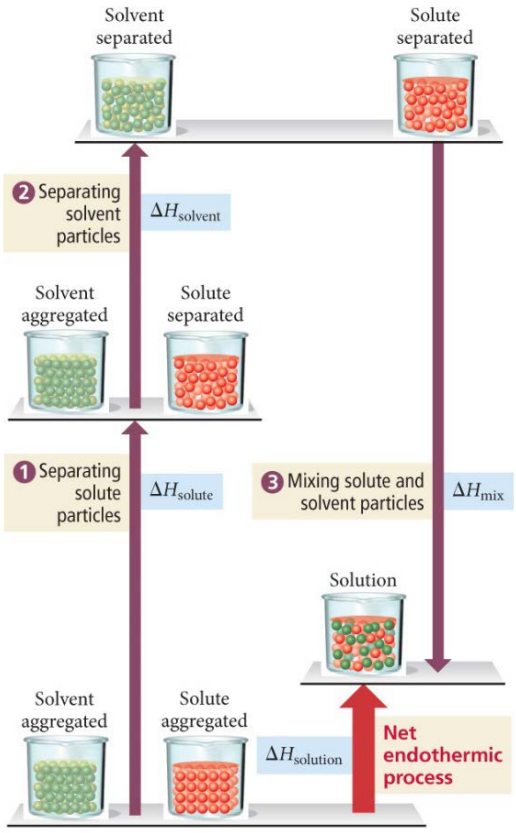


Soluble or Insoluble?

Soluble	Insoluble
Higher solute-solvent attractions	Lower solute-solvent attractions
	
Soluble (miscible)	Insoluble (immiscible)

Notice, in general, the solute particles that are **INSOLUBLE** have _____ ion charges, which means they have a greater attraction to other solute particles: go, Coulomb's Law, go!

The Thermodynamics of Dissolution

Exothermic Dissolution ($-\Delta H_{soln}$)	**Endothermic Dissolution** ($+\Delta H_{soln}$)
• Heat released when salt dissolve • Feels warm to the touch	• Heat absorbed when salt dissolve • Feels cold to the touch
<u>Always</u> thermodynamically favorable ($-\Delta G$) because entropy will always increase $-\Delta H, +\Delta S$	**Thermodynamically favorable at warmer temperatures**, depending on increase in entropy $+\Delta H, +\Delta S$
	

Handy rule of thumb: **"Like dissolves like"**

Great multiple choice trick but does _____ count as explanation on free response!

1. Polar solvents dissolve polar solutes
    - Hydrophilic (polar) groups to watch for: OH, CHO, C=O, COOH, NH_2, and Cl.
2. Non-polar solvents dissolve non-polar solutes
    - Hydrophobic (non-polar) groups to watch for: C – H and C – C.

Many molecules have both hydrophilic and hydrophobic parts; solubility in water becomes a _____ between the attraction of the polar groups for water and the attraction of the nonpolar groups for their own kind.

Never use "like dissolves like" to explain a FR on the AP exam: instead, EXPLAIN in terms of structure, IMFs, and energy!

So… how *do* you explain solubility for free response questions?

1. **Identify solute-solvent IMFS**

Type of Substance	Dominant Interaction with Water	Dominant Interaction with a Non-polar Solvent
Ionic	ion-dipole	ion-induced dipole (nope)
Polar + FON	hydrogen bonds	dipole-induced dipole
Polar	dipole-dipole	dipole-induced dipole
Non-polar	dipole-induced dipole	induced dipole-induced dipole (London dispersion forces)

2. **Are solute–solvent attractions _____ than solute–solute (or solvent–solvent) attractions?**
    - Explain: strong interactions BETWEEN solvent and solute → yes, solute will dissolve!
3. **Solute–solvent attractions _____ than solute–solute (or solvent–solvent) attractions?**
    - Explain: weak solute–solvent interactions are not as strong as existing solvent–solvent (or solute-solute) attractions, thus solute will _____ dissolve.

Of course, you must be _____! Identify **BOTH** solute and solvent by name or formula.

*Note: you do NOT have to explain WHY a given compound can form specific IMFS; it is enough to state them.

Example #1: Can CH_3OH dissolve in water? Why or why not?

Too much "CH_3OH can form hydrogen bonds with water because it has a hydrogen which is covalently bonded to an oxygen, so it will form strong IMFs with water and thus will be able to dissolve in water."

Just right "CH_3OH can form strong hydrogen bonds with water, so it will be able to dissolve."

Not enough "CH_3OH can form strong hydrogen bonds, so it will be able to dissolve in water."

Example #2: Can benzene, C_6H_6, dissolve in water? Why or why not?

Too much "C_6H_6 is non-polar with has a dipole moment of zero, and so it can only form weak dipole-induced dipole interactions with water, which are not as strong as the hydrogen bonds that already exist between water molecules, so C_6H_6 won't dissolve in water."

Just right "C_6H_6 is non-polar and can only form weak intermolecular attractions with water, which are not as strong as the hydrogen bonds that already exist between water molecules, so C_6H_6 won't dissolve in water."

Not enough "C_6H_6 is non-polar, so it won't dissolve in a polar substance like water."

Example FR question: Which is more likely to be soluble in water, liquid methanol (CH_3OH) or liquid hexane (C_6H_{14})? Justify your answer.

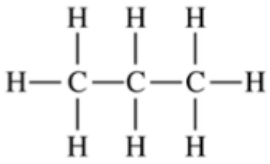
Free Response Practice!

Directions: Use principles of atomic structure, bonding, and intermolecular forces to answer the following questions. Your responses must include specific information about all substances referred to in each part.

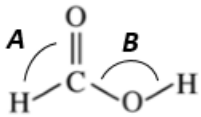
1. Ammonia, NH_3, is very soluble in water, whereas phosphine, PH_3, is only moderately soluble in water. Explain.

2. Indicate whether you agree or disagree with the statement in the box below. Justify your answer.

H_2O has a much higher boiling point than H_2S because it has very strong hydrogen bonding forces between its molecules as well as dipole−dipole forces and London dispersion forces, while H_2S has only dipole-dipole and dispersion forces. The stronger hydrogen bonding IMFs between water molecules mean that the bonds between water molecules are harder to break.



Propane



Methanoic Acid

3. The complete structural formulas of propane, C_3H_8, and methanoic acid, HCOOH, are shown above.
   a. In the table below, write the type(s) of intermolecular attractive force(s) that occur in each substance.

Substance	Boiling Point	Intermolecular Attractive Force(s)
Propane	229 K	
Methanoic acid	374 K	

   b. Explain why methanoic acid has a higher boiling point than propane.

   c. Which bond angle would you predict to be larger: angle A (H – C – O) or angle B (C – O – H) in methanoic acid? Justify your answer.

Bonding and IMFS Free Response Study Guide

1. Draw the dang Lewis dot structure!
2. Reference the chart below depending on the question type.

Question Type	Example Question(s)	What to consider/talk about:
Expanded octet exceptions	• Which atoms can be stable with more than 8 bonded valence electrons and why?	• Must have access to <u>empty d sublevel</u> • Only available to elements in rows 3-7
Comparing bond length and/or strength	• Which bond is shorter, X_2 or Y_2? • Which bond is stronger, X_2 or Y_2?	More shared electron pairs means: o Stronger/shorter bond o Greater attractive force between e^- and nuclei o Nuclei pulled closer together/ harder to separate
Comparing bond angles	• Which bond angle is larger/smaller, H – X – H or H – Y – H?	• Count number of lone pairs vs bonding pairs on central atom • <u>Lone pairs are more repulsive than bonding pairs</u> • More lone pairs = smaller bond angle
Comparing VSEPR geometry	• Why does XH_2 have bent geometry, but YH_2 is linear? • Why does XH_3 have trigonal planar geometry, but YH_3 is trigonal pyramidal?	• Count number of lone pairs vs bonding pairs on central atom • Lone pairs distort the symmetry, pushing bonded atoms away
Comparing polar vs non-polar <u>bonds</u>	• Which bond is more polar, HX or HY? • The X_2 bond is non-polar. Explain.	• Greater electronegativity difference between bonded atoms = <u>more uneven distribution of e^- density</u> = more polar • Non-polar = equally distributed e^- density
Comparing polar vs non-polar <u>molecules</u>	• Why is XH_3 a polar molecule, but YH_3 is non-polar?	• Lone pair on central atom distorts symmetry = <u>uneven distribution of e^- density</u> = polar • No lone pair on central atom = symmetrical e^- density = non-polar
Comparing vapor pressures, boiling points, or melting points	• Why does XH_2 have a lower vapor pressure than YH_2? • Which has a higher melting point, XH_3 or YH_3? Explain.	• Identify IMFS o Non-polar = LDFs o Polar = dipole to dipole o H – FON = hydrogen bonding • Connect IMF strength to vocab o ↑ IMFs = ↓ VP, because molecules are more attracted to each other, less in gaseous state o ↑ IMFs = ↑ BP/MP, because molecules are more attracted to each other and require more energy to separate

But WAIT! **How can I tell if I molecule is polar?**

<table>
<tr><td rowspan="3">Lone pairs on the central atom?</td><td>Yes</td><td colspan="2">Polar molecule</td></tr>
<tr><td rowspan="2">No</td><td colspan="2">Look at terminal (non-central atoms)</td></tr>
<tr><td>Same?</td><td>Different?</td></tr>
<tr><td>Non-polar molecule</td><td>Polar molecule</td></tr>
</table>

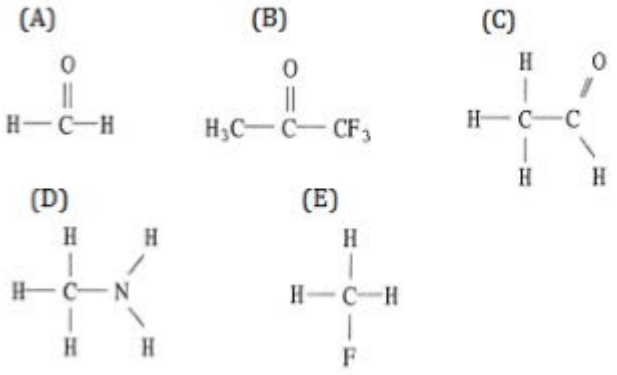
Unit 11 Multiple Choice Practice

Hydrogen Halide	Normal Boiling Point, °C
HF	+19
HCl	– 85
HBr	– 67
HI	– 35

1. The liquified hydrogen halides have the normal boiling points given above. The relatively high boiling point of HF can be correctly explained by which of the following?
   a. HF molecules tend to form hydrogen bonds.
   b. HF is the strongest acid.
   c. HF molecules have a smaller dipole moment.
   d. HF is much less soluble in water.

2. Which of the following compounds would have the highest lattice energy?
   a. LiF b. $MgCl_2$ c. $CaBr_2$ d. C_2H_6

3. Which one of the following substances will have hydrogen bonding as one of its intermolecular forces?



4. The substance with the lowest boiling point is _____ .
   a. I_2 b. Br_2 c. Cl_2 d. F_2

5. Which of the following substances is an electrolyte when dissolved in water?
   a. CH_3COOH b. HCl c. AsF_3 d. SeF_2

6. Which of the following pairs of elements is most likely to create an interstitial alloy?
   a. titanium and copper
   b. aluminum and lead
   c. silver and tin
   d. magnesium and calcium

7. The relationship of the following two structures is

a. resonance structures
b. isomers
c. both isomers and resonance structures
d. neither isomers nor resonance structures

8. Given the following substances and their boiling points:

C: 43.8 °C D: 93.7 °C M: 56.7 °C T: 83.5 °C R: 63.6 °C

Which ranking correctly lists some of these substances in order of decreasing intermolecular forces?

a. C > R > D
b. D > T > R
c. R > M > D
d. C > D > M

9. When NaCl dissolves in water, aqueous Na^+ and Cl^- result. The force of attraction that exists between Na^+ and H_2O is called a(n) _____ interaction.

a. Dipole – dipole
b. Ion – ion
c. Hydrogen bonding
d. Ion – dipole
e. London dispersion force

10. Why does CaF_2 have a higher melting point than NH_3?

a. CaF_2 is more massive and thus has stronger London dispersion forces.
b. CaF_2 exhibits network covalent bonding, which is the strongest type of bonding.
c. CaF_2 is smaller and exhibits greater Coulombic attractive forces.
d. CaF_2 is an ionic substance and it requires a lot of energy to break up an ionic substance.

11. If X represents the central atom, which molecule has the largest F–X–F bond angle?

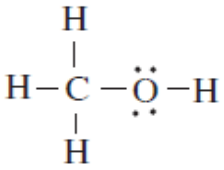
a. OF_2 b. CF_4 c. BF_3 d. NF_3

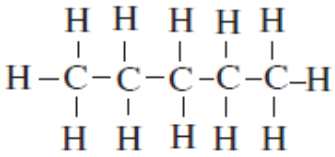
12. Why can a molecule with the structure NBr_5 not exist?

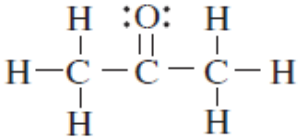
a. Nitrogen only has two energy levels and thus is unable to expand its octet.
b. Bromine is much larger than nitrogen and cannot be a terminal atom in this molecule.
c. It is impossible to complete the octet for all six atoms using only valence electrons.
d. Nitrogen does not have a low enough electronegativity to be the central atom of this molecule.

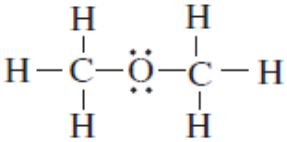
13. A sample of liquid NH_3 is brought to its boiling point. Which of the following occurs during the boiling process?
    a. The N−H bonds within the NH_3 molecules break apart.
    b. The overall temperature of the solution rises as the NH_3 molecules speed up.
    c. The amount of energy within the system remains constant.
    d. The hydrogen bonds holding separate NH_3 molecules together break apart.

14. The following diagrams show the Lewis structures of four different molecules. Which molecule would be most soluble in a non-polar solvent?


Methanol


Pentane


Acetone


Ether

    a. methanol b. pentane c. acetone d. ether

15. Which Lewis structure indicates the presence of two lone pairs of electrons on the central atom?
    a. BeF_2 b. NH_3 c. OCl_2 d. CH_2Cl_2

16. Which bond angle is present in the ammonium ion, NH_4^+?
    a. 90° b. 180° c. 120° d. 109.5°

17. Which molecule is non-polar?
    a. BCl_3 b. NCl_3 c. PCl_3 d. $CHCl_3$

18. Predict the shape of of a molecule of nitrogen trifluoride.
    a. bent b. trigonal planar
    b. tetrahedral d. trigonal pyramidal

19. Which molecule is polar?

a. BCl_3 b. NCl_3 c. PCl_5 d. CCl_4

20. Use VSEPR theory to predict the shape of carbon dioxide, CO_2.

a. bent b. trigonal planar

b. tetrahedral d. linear

21. What is the hybridization of the carbon atom in carbon monoxide, CO?

a. sp b. sp^2 c. sp^3 b. sp^3d

22. All of the following molecules contain polar bonds. Of these molecules, the <u>only one</u> that is a non-polar molecule is:

a. H_2O b. CO_2 c. HCl d. NH_3

23. Which of the following compounds contains at least one element that violates the octet rule?

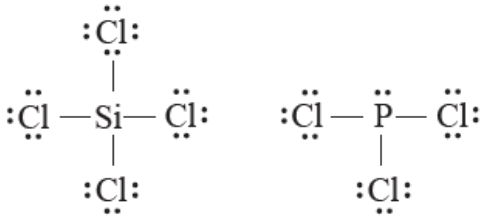
a. OF_2 b. NF_3 c. CO_2 d. H_2O

24. The six carbon atoms in a benzene molecule are shown in different resonance forms as three single bonds and three double bonds. If the length of a single carbon−carbon bond is 154 pm and the length of a double carbon−carbon bond is 133 pm, what length would be expected for the carbon−carbon bonds in benzene?

a. 126 pm b. 133 pm c. 140 pm d. 154 pm

25. London dispersion forces are caused by

a. temporary dipoles created by the position of electrons around the nuclei in a molecule
b. the three-dimensional intermolecular bonding present in all covalent substances
c. the uneven electron-to-proton ratio found on the individual atoms of a molecule
d. the electronegativity differences between different atoms in a molecule

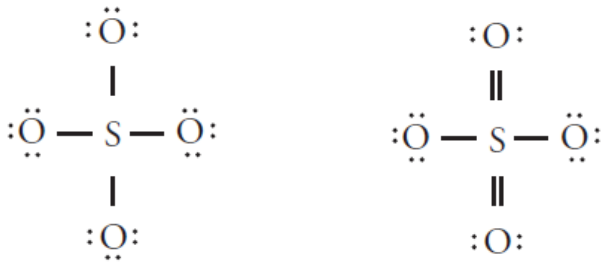


26. The Lewis diagrams for $SiCl_4$ and PCl_3 are drawn above. What are the approximate bond angles between the terminal chlorine atoms in each structure?

	$SiCl_4$	PCl_3
(A)	90°	90°
(B)	109.5°	< 109.5°
(C)	90°	109.5°
(D)	< 109.5°	> 90°

Use the following Lewis diagrams to answer questions 27–29.

There are several potential different Lewis electron-dot structures for the sulfate ion, two of which are below.



Structure A Structure B

27. What is the molecular geometry in the structure A?

a. octahedral b. trigonal planar

b. tetrahedral d. trigonal pyramidal

28. What is the S–O bond order in structure B?

a. 1.0 b. 1.33 c. 1.5 d. 1.67

29. Which structure is more likely to correspond with the actual Lewis diagram for the sulfate ion?

a. Structure A; single bonds are more stable than double bonds

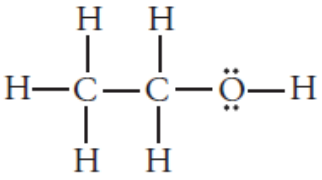
b. Structure A; it has the most unshared pairs of electrons

c. Structure B; there are more possible resonance structures

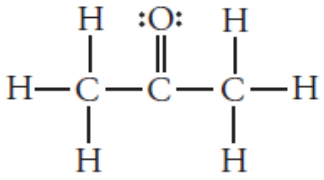
d. Structure B; fewer atoms have formal charges

Use the following Lewis diagrams to answer questions 30–32.

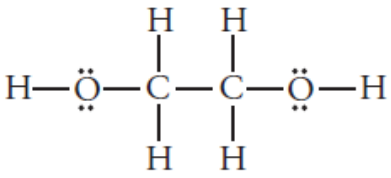
The following three substances are kept in identical containers at 25°C. All three are in the liquid phase.



Ethanol



Acetone



Ethylene Glycol

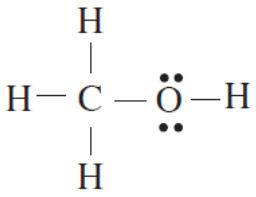
30. Which substance would have the highest boiling point?
    a. Ethanol, because it is most asymmetrical.
    b. Acetone, because of the double bond.
    c. Ethylene glycol, because it has the most hydrogen bonding.
    d. All three substance would have very similar boiling points because their molar masses are similar.

31. Which substance would have the highest vapor pressure?
    a. Ethanol, because of the hybridization of its carbon atoms.
    b. Acetone, because it exhibits the weakest intermolecular forces.
    c. Ethylene glycol, because it has the most lone pairs assigned to individual atoms.
    d. All three substance would have very similar vapor pressure because they have a similar number of electrons.

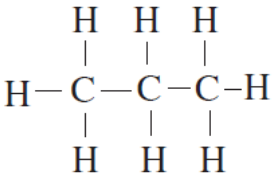
32. Which of the substances would be soluble in water?
    a. Ethylene glycol only, because it has the longest bond lengths.
    b. Acetone only, because it is the most symmetrical.
    c. Ethanol and ethylene glycol, because of their hydroxyl (−OH) group
    d. All three substance would be soluble in water due to their permanent dipoles.

33. Which of the following has a dipole moment of zero?
    a. PF_5 b. HCN c. SO_2 d. NH_3

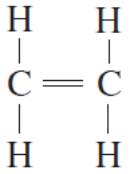
Questions 34–36 refer to the following structures.



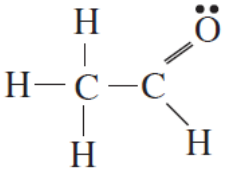
Methanol



Propane



Ethene



Ethanal

34. Based on the strength of the intermolecular forces in each substance estimate from greatest to smallest the vapor pressures of each substance in liquid state at the same temperature.
    a. Propane > Ethanal > Ethene > Methanol
    b. Ethene > Propane > Ethanal > Methanol
    c. Ethanal > Methanol > Ethene > Propane
    d. Methanol > Ethanal > Propane > Ethene

35. When in liquid state, which two substances are most likely to be miscible with water?
    a. Propane and ethene
    b. Methanol and propane
    c. Ethene and ethanal
    d. Methanol and ethanal

36. Between propane and ethene, which will likely have the higher boiling point and why?
    a. Propane, because it has a greater molar mass
    b. Propane, because it has a more polarizable electron cloud.
    c. Ethene, because of the double bond.
    d. Ethene, because it is smaller in size.

37. The bond length between any two nonmetal atoms is achieved under which of the following conditions?
    a. Where the energy of interaction between the atoms is at its minimum value
    b. Where the nuclei of each atom exhibit the strongest attraction to the electrons of the other atom
    c. The point at which the attractive and repulsive forces between the two atoms are equal
    d. The closest point at which a valence electron from one atom can transfer to the other atom.

Use the following information to answer questions 38–41.

Consider the Lewis structures for the following molecules:

CO_2, CO_3^{2-}, NO_2^-, and NO_3^-

38. Which molecule would have the shortest bonds?

a. CO_3^{2-} b. NO_2^- c. CO_2 d. NO_3^-

39. Which molecules are best represented by multiple resonance structures?

a. CO_3^{2-} and NO_3^- c. CO_2 and CO_3^{2-}

b. CO_3^{2-}, NO_2^- and NO_3^- d. NO_2^- and NO_3^-

40. Which molecule or molecules exhibit sp^2 hybridization around the central atom?

a. CO_3^{2-} and NO_3^- c. CO_2 and CO_3^{2-}

b. CO_3^{2-}, NO_2^- and NO_3^- d. NO_2^- and NO_3^-

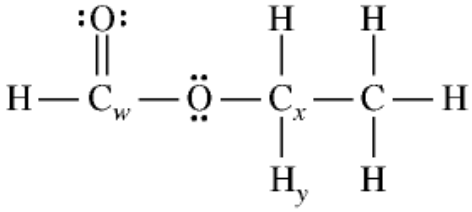
41. Which molecule would have the smallest bond angle between terminal atoms?

a. CO_3^{2-} b. NO_2^- c. CO_2 d. NO_3^-

FR Practice #1 (2011B #6 shortened, 4 points)

1. Use principles of molecular structure, intermolecular forces, and kinetic molecular theory to answer the following questions.

   a. A complete Lewis electron-dot diagram of a molecule of ethyl methanoate is given below.



      i. Identify the hybridization of the valence electrons of the carbon atom labeled C_w. (1 point)

      ii. Estimate the numerical value of the H_y–C_x–O bond angle in an ethyl methanoate molecule. Explain the basis of your estimate. (1 point)

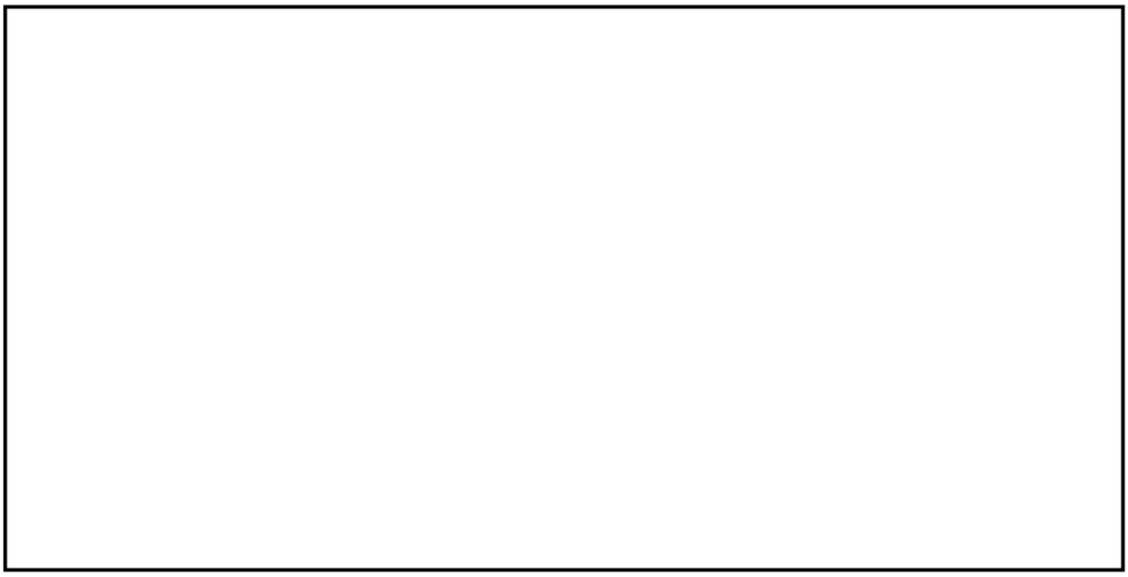
   b. Ethyl methanoate, CH_3CH_2OCHO, is synthesized in the laboratory from ethanol, C_2H_5OH, and methanoic acid, HCOOH, as represented by the following equation.

$$C_2H_5OH(l) + HCOOH(l) \rightleftarrows CH_3CH_2OCHO(l) + H_2O(l)$$

      i. In the box below, draw the complete Lewis electron-dot diagram of a methanoic acid molecule. (1 point)

Methanoic Acid

ii. In the box below, draw the complete Lewis electron-dot diagrams of a methanoic acid molecule and a water molecule in an orientation that allows a hydrogen bond to form between them. (1 point)

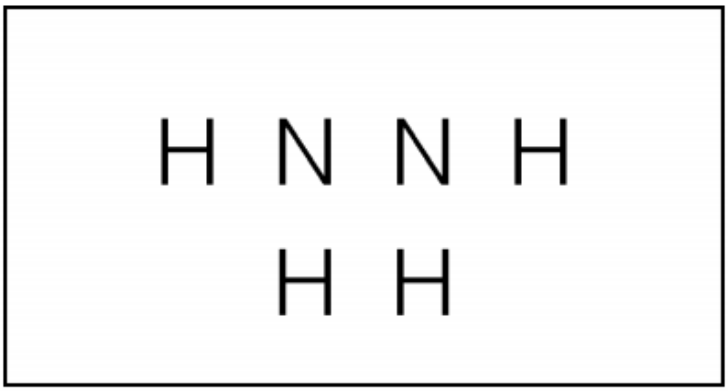


Hydrogen Bonding Between Methanoic Acid and Water

FR Practice #2 (2011 #5, 8 points)

2. Hydrazine is an inorganic compound with the formula N_2H_4.

   a. In the box below, complete the Lewis electron-dot diagram for the N_2H_4 molecule by drawing in all the electron pairs. (1 point)



   b. On the basis of the diagram you completed in part (a), do all six atoms in the N_2H_4 molecule lie in the same plane? Explain. (1 point)

   c. The normal boiling point of N_2H_4 is 114°C, whereas the normal boiling point of C_2H_6 is −89°C. Explain, in terms of the intermolecular forces present in each liquid, why the boiling point of N_2H_4 is so much higher than that of C_2H_6. (2 points)

   d. Write a balanced chemical equation for the reaction between N_2H_4 and H_2O that explains why a solution of hydrazine in water has a pH greater than 7. (1 point)

N_2H_4 reacts in air according to the equation below.

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2\ H_2O(g) \qquad \Delta H^\circ = -534\ \text{kJ mol}^{-1}$$

e. Is the reaction an oxidation-reduction, acid-base, or decomposition reaction? Justify your answer. (1 point)

f. Predict the sign of the entropy change, ΔS, for the reaction. Justify your prediction. (1 point)

g. Indicate whether the statement written in the box below is true or false. Justify your answer. (1 point)

The large negative ΔH° for the combustion of hydrazine results from the large release of energy that occurs when the strong bonds of the reactants are broken.

FR Practice #3 (2016 #1, shortened, 6 points)

3. A student investigates the enthalpy of solution, ΔH_{soln}, for two alkali metal halides, LiCl and NaCl. To explain why ΔH_{soln} for NaCl is different than that for LiCl, the student investigates factors that affect ΔH_{soln} and finds that ionic radius and lattice enthalpy (which can be defined as the ΔH associated with the separation of a solid crystal into gaseous ions) contribute to the process. The student consults references and collects the data shown in the table below.

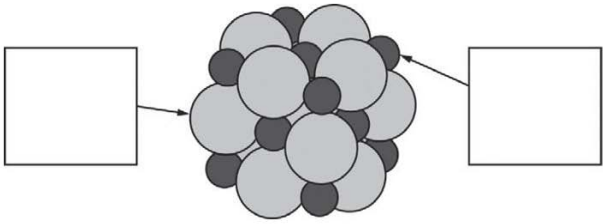
Ion	Ionic Radius (pm)
Li^+	76
Na^+	102

a. Write the complete electron configuration for the Na^+ ion in the ground state. (1 point)

b. Using principles of atomic structure, explain why the Na^+ ion is larger than the Li^+ ion. (1 point)

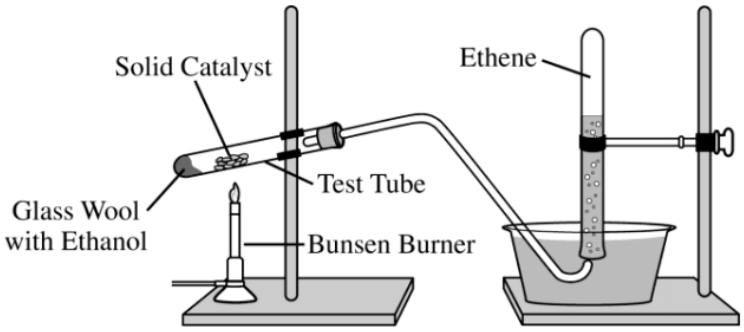
c. Which salt, LiCl or NaCl, has the greater lattice enthalpy? Justify your answer. (1 point)

d. Below is a representation of a portion of a crystal of LiCl. Identify the ions in the representation by writing the appropriate formulas (Li^+ or Cl^-) in the boxes below. (1 point)



e. The lattice enthalpy of LiCl is positive, indicating that it takes energy to break the ions apart in LiCl. However, the dissolution of LiCl in water is an exothermic process. Identify all particle-particle interactions that contribute significantly to the dissolution process being exothermic. For each interaction, include the particles that interact and the specific type of intermolecular force between those particles. (2 points)

FR Practice #4 (2015 #2 (shortened), 4 points)

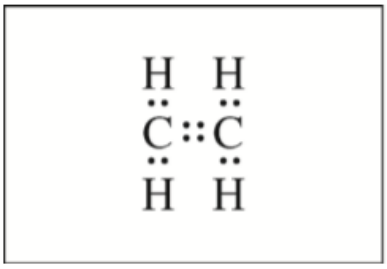


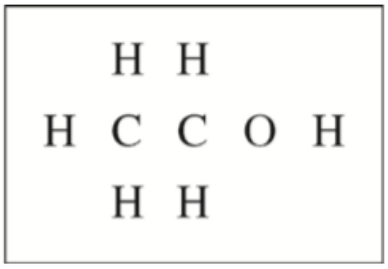
4. Ethene, $C_2H_4(g)$ (molar mass 28.1 g/mol), may be prepared by the dehydration of ethanol, $C_2H_5OH(g)$ (molar mass 46.1 g/mol), using a solid catalyst. A setup for the lab synthesis is shown in the diagram above. The equation for the dehydration reaction is given below.

$$\underset{\text{ethanol}}{C_2H_5OH(g)} \xrightarrow{\text{catalyst}} \underset{\text{ethene}}{C_2H_4(g)} + \underset{\text{water}}{H_2O(g)} \quad \Delta H^\circ_{298} = 45.5 \text{ kJ/mol}_{rxn};\ \Delta S^\circ_{298} = 126 \text{ J/(K·mol}_{rxn})$$

A student added a 0.200 g sample of $C_2H_5OH(l)$ to a test tube using the setup shown above. The student heated the test tube gently with a Bunsen burner until all of the $C_2H_5OH(l)$ evaporated and gas generation stopped.

a. The Lewis electron-dot diagram for C_2H_4 is shown below in the box on the left. In the box on the right, complete the Lewis electron-dot diagram for C_2H_5OH by drawing in all of the electron pairs. (1 point)





b. What is the approximate value of the H – C – H bond angle in the ethene molecule? Explain. (1 point)

c. During the dehydration experiment, $C_2H_4(g)$ and unreacted $C_2H_5OH(g)$ passed through the tube into the water. The C_2H_4 was quantitatively collected as a gas, but the unreacted C_2H_5OH was not. Explain this observation in terms of the intermolecular forces between water and each of the two gases. (2 points)

Free Response Practice #5

1. Consider the four covalent compound $GeCl_4$, $SeCl_4$, ICl_4^-, and ICl_4^+.

   a. Draw the Lewis structure (electron-dot diagram) of each of the four covalent species in the boxes provided below. Show all valence electrons in your structures. (4 points)

$GeCl_4$	$SeCl_4$
ICl_4^-	ICl_4^+

   b. On the basis of the Lewis structures drawn in part (a), answer the following questions.

      i. Identify the Cl–I–Cl bond angle in ICl_4^-.

      ii. What is the hybridization of the Ge atom in $GeCl_4$?

      iii. What is the geometric shape formed by the atoms in ICl_4^+?

      iv. How many sigma bonds and how many pi bonds are in the ICl_4^- structure?

      v. Is $SeCl_4$ polar? Explain.

Free Response Practice #6 (2017 #2 (shortened) and 2018 #2 (shortened))

1. Answer the following question about fulminic acid, HCNO.

   Two possible Lewis electron-dot diagrams for fulminic acid are shown below.

   Explain why the diagram on the left is the better representation for the bonding in fulminic acid. Justify your choice based on formal charges.

2. $N_2O_3(g)$ reacts with water to form nitrous acid, $HNO_2(aq)$, a compound involved in the production of acid rain. The reaction is represented below.

   The skeletal structure of the HNO_2 molecule is shown in the box below.

   a. Complete the Lewis electron-dot diagram of the HNO_2 molecule in the box below, including any lone pairs of electrons.

   b. Based on your completed diagram, identify the hybridization of the nitrogen atom in the HNO_2 molecule.

   c. Identify the O–N–O bond angle in nitrous acid.

STANDARD REDUCTION POTENTIALS IN AQUEOUS SOLUTION AT 25°C

Half-reaction			$E°$(V)
$F_2(g) + 2\,e^-$	→	$2\,F^-$	2.87
$Co^{3+} + e^-$	→	Co^{2+}	1.82
$Au^{3+} + 3\,e^-$	→	$Au(s)$	1.50
$Cl_2(g) + 2\,e^-$	→	$2\,Cl^-$	1.36
$O_2(g) + 4\,H^+ + 4\,e^-$	→	$2\,H_2O(l)$	1.23
$Br_2(l) + 2\,e^-$	→	$2\,Br^-$	1.07
$2\,Hg^{2+} + 2\,e^-$	→	Hg_2^{2+}	0.92
$Hg^{2+} + 2\,e^-$	→	$Hg(l)$	0.85
$Ag^+ + e^-$	→	$Ag(s)$	0.80
$Hg_2^{2+} + 2\,e^-$	→	$2\,Hg(l)$	0.79
$Fe^{3+} + e^-$	→	Fe^{2+}	0.77
$I_2(s) + 2\,e^-$	→	$2\,I^-$	0.53
$Cu^+ + e^-$	→	$Cu(s)$	0.52
$Cu^{2+} + 2\,e^-$	→	$Cu(s)$	0.34
$Cu^{2+} + e^-$	→	Cu^+	0.15
$Sn^{4+} + 2\,e^-$	→	Sn^{2+}	0.15
$S(s) + 2\,H^+ + 2\,e^-$	→	$H_2S(g)$	0.14
$2\,H^+ + 2\,e^-$	→	$H_2(g)$	0.00
$Pb^{2+} + 2\,e^-$	→	$Pb(s)$	−0.13
$Sn^{2+} + 2\,e^-$	→	$Sn(s)$	−0.14
$Ni^{2+} + 2\,e^-$	→	$Ni(s)$	−0.25
$Co^{2+} + 2\,e^-$	→	$Co(s)$	−0.28
$Cd^{2+} + 2\,e^-$	→	$Cd(s)$	−0.40
$Cr^{3+} + e^-$	→	Cr^{2+}	−0.41
$Fe^{2+} + 2\,e^-$	→	$Fe(s)$	−0.44
$Cr^{3+} + 3\,e^-$	→	$Cr(s)$	−0.74
$Zn^{2+} + 2\,e^-$	→	$Zn(s)$	−0.76
$2\,H_2O(l) + 2\,e^-$	→	$H_2(g) + 2\,OH^-$	−0.83
$Mn^{2+} + 2\,e^-$	→	$Mn(s)$	−1.18
$Al^{3+} + 3\,e^-$	→	$Al(s)$	−1.66
$Be^{2+} + 2\,e^-$	→	$Be(s)$	−1.70
$Mg^{2+} + 2\,e^-$	→	$Mg(s)$	−2.37
$Na^+ + e^-$	→	$Na(s)$	−2.71
$Ca^{2+} + 2\,e^-$	→	$Ca(s)$	−2.87
$Sr^{2+} + 2\,e^-$	→	$Sr(s)$	−2.89
$Ba^{2+} + 2\,e^-$	→	$Ba(s)$	−2.90
$Rb^+ + e^-$	→	$Rb(s)$	−2.92
$K^+ + e^-$	→	$K(s)$	−2.92
$Cs^+ + e^-$	→	$Cs(s)$	−2.92
$Li^+ + e^-$	→	$Li(s)$	−3.05

Polyatomic Ions

1st six-weeks

Nick the **C**amel ***ate*** an **I**cky **Cl**am for **S**upper in **P**hoenix with his **Br**os

NO_3^- nitrate	ClO_3^- chlorate	PO_3^{3-} phosphite
NO_2^- nitrite	ClO_2^- chlorite	BrO_3^- bromate
CO_3^{2-} carbonate	SO_4^{2-} sulfate	BrO_2^- bromite
IO_3^- iodate	SO_3^{2-} sulfite	
IO_2^- iodite	PO_4^{3-} phosphate	

Consonants = # of Oxygen

Vowels = Charge

2nd six-weeks

ClO_4^- perchlorate	IO^- hypoiodite
ClO^- hypochlorite	BrO_4^- perbromate
IO_4^- periodate	BrO^- hypobromite

	Difference in Oxygen from ATE
Per_____ate	+1
Ate	0
Ite	−1
Hypo ____ite	−2

3rd six-weeks

$H_2PO_4^-$ dihydrogen phosphate	HCO_3^- hydrogen carbonate *or* bicarbonate
HPO_4^{2-} hydrogen phosphate	HSO_4^- hydrogen sulfate

4th six-weeks

NH_4^+ ammonium	OH^- hydroxide
$C_2H_3O_2^-$ *or* CH_3COO^- acetate	H_3O^+ hydronium

5th six-weeks

MnO_4^- permanganate	CrO_4^{2-} chromate
CN^- cyanide	$Cr_2O_7^{2-}$ dichromate

6th six-weeks

O_2^{2-} peroxide	$C_4H_4O_6^{2-}$ tartrate
$S_2O_3^{2-}$ thiosulfate	$C_2O_4^{2-}$ oxalate

Acid Nomenclature	
Binary	Hydro_____ic
ate	ic
ite	ous

Polyatomic Ions

Br–Based Ions

BrO^-	hypobromite
BrO_2^-	bromite
BrO_3^-	bromate
BrO_4^-	perbromate

Cr–Based Ions

CrO_4^{2-}	chromate
$Cr_2O_7^{2-}$	dichromate

I–Based Ions

IO^-	hypoiodite
IO_2^-	iodite
IO_3^-	iodate
IO_4^-	periodate

P–Based Ions

PO_3^{3-}	phosphite
PO_4^{3-}	phosphate
HPO_4^{2-}	hydrogen phosphate
$H_2PO_4^-$	dihydrogen phosphate

Other Ions

CN^-	cyanide
O_2^{2-}	peroxide
MnO_4^-	permanganate

C–Based Ions

CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate *or* bicarbonate
$C_2H_3O_2^-$ *or* CH_3COO^-	acetate
$C_4H_4O_6^{2-}$	tartrate
$C_2O_4^{2-}$	oxalate

Cl–Based Ions

ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate

N–Based Ions

NO_2^-	nitrite
NO_3^-	nitrate
NH_4^+	ammonium

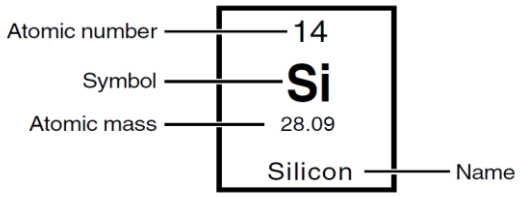
S–Based Ions

SO_3^{2-}	sulfite
SO_4^{2-}	sulfate
HSO_4^-	hydrogen sulfate
$S_2O_3^{2-}$	thiosulfate

Acid & Base Ions

H_3O^+	hydronium
OH^-	hydroxide

PERIODIC TABLE OF THE ELEMENTS



	1																8
1	1 **H** 1.008 Hydrogen	2										3	4	5	6	7	2 **He** 4.00 Helium
2	3 **Li** 6.94 Lithium	4 **Be** 9.01 Beryllium										5 **B** 10.81 Boron	6 **C** 12.01 Carbon	7 **N** 14.01 Nitrogen	8 **O** 16.00 Oxygen	9 **F** 19.00 Fluorine	10 **Ne** 20.18 Neon
3	11 **Na** 22.99 Sodium	12 **Mg** 24.30 Magnesium										13 **Al** 26.98 Aluminum	14 **Si** 28.09 Silicon	15 **P** 30.97 Phosphorus	16 **S** 32.06 Sulfur	17 **Cl** 35.45 Chlorine	18 **Ar** 39.95 Argon
4	19 **K** 39.10 Potassium	20 **Ca** 40.08 Calcium	21 **Sc** 44.96 Scandium	22 **Ti** 47.90 Titanium	23 **V** 50.94 Vanadium	24 **Cr** 52.00 Chromium	25 **Mn** 54.94 Manganese	26 **Fe** 55.85 Iron	27 **Co** 58.93 Cobalt	28 **Ni** 58.69 Nickel	29 **Cu** 63.55 Copper	30 **Zn** 65.39 Zinc	31 **Ga** 69.72 Gallium	32 **Ge** 72.59 Germanium	33 **As** 74.92 Arsenic	34 **Se** 78.96 Selenium	35 **Br** 79.90 Bromine	36 **Kr** 83.80 Krypton
5	37 **Rb** 85.47 Rubidium	38 **Sr** 87.62 Strontium	39 **Y** 88.91 Yttrium	40 **Zr** 91.22 Zirconium	41 **Nb** 92.91 Niobium	42 **Mo** 95.94 Molybdenum	43 **Tc** (98) Technetium	44 **Ru** 101.10 Ruthenium	45 **Rh** 102.91 Rhodium	46 **Pd** 106.42 Palladium	47 **Ag** 107.87 Silver	48 **Cd** 112.41 Cadmium	49 **In** 114.82 Indium	50 **Sn** 118.71 Tin	51 **Sb** 121.75 Antimony	52 **Te** 127.60 Tellurium	53 **I** 126.91 Iodine	54 **Xe** 131.29 Xenon
6	55 **Cs** 132.91 Cesium	56 **Ba** 137.33 Barium	57 ***La** 138.91 Lanthanum	72 **Hf** 178.49 Hafnium	73 **Ta** 180.95 Tantalum	74 **W** 183.85 Tungsten	75 **Re** 186.21 Rhenium	76 **Os** 190.2 Osmium	77 **Ir** 192.2 Iridium	78 **Pt** 195.08 Platinum	79 **Au** 196.97 Gold	80 **Hg** 200.59 Mercury	81 **Tl** 204.38 Thallium	82 **Pb** 207.2 Lead	83 **Bi** 208.98 Bismuth	84 **Po** (209) Polonium	85 **At** (210) Astatine	86 **Rn** (222) Radon
7	87 **Fr** (223) Francium	88 **Ra** 226.02 Radium	89 **†Ac** 227.03 Actinium	104 **Rf** (261) Rutherfordium	105 **Db** (262) Dubnium	106 **Sg** (266) Seaborgium	107 **Bh** (264) Bohrium	108 **Hs** (277) Hassium	109 **Mt** (268) Meitnerium	110 **Ds** (271) Darmstadtium	111 **Rg** (272) Roentgenium							

***Lanthanide Series**	58 **Ce** 140.12 Cerium	59 **Pr** 140.91 Praseodymium	60 **Nd** 144.24 Neodymium	61 **Pm** (145) Promethium	62 **Sm** 150.4 Samarium	63 **Eu** 151.97 Europium	64 **Gd** 157.25 Gadolinium	65 **Tb** 158.93 Terbium	66 **Dy** 162.5 Dysprosium	67 **Ho** 164.93 Holmium	68 **Er** 167.26 Erbium	69 **Tm** 168.93 Thulium	70 **Yb** 173.04 Ytterbium	71 **Lu** 174.97 Lutetium
†Actinide Series	90 **Th** 232.04 Thorium	91 **Pa** 231.04 Protactnium	92 **U** 238.03 Uranium	93 **Np** (237) Neptumium	94 **Pu** (244) Plutonium	95 **Am** (243) Americium	96 **Cm** (247) Curium	97 **Bk** (247) Berkelium	98 **Cf** (251) Californium	99 **Es** (252) Einsteinium	100 **Fm** (257) Fermium	101 **Md** (258) Mendelevium	102 **No** (259) Nobelium	103 **Lr** (262) Lawrencium

AP® CHEMISTRY EQUATIONS AND CONSTANTS

Throughout the exam the following symbols have the definitions specified unless otherwise noted.

L, mL	=	liter(s), milliliter(s)	mm Hg	=	millimeters of mercury
g	=	gram(s)	J, kJ	=	joule(s), kilojoule(s)
nm	=	nanometer(s)	V	=	volt(s)
atm	=	atmosphere(s)	mol	=	mole(s)

ATOMIC STRUCTURE

$E = h\nu$

$c = \lambda\nu$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s
Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number = 6.022×10^{23} mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$, where $a\,A + b\,B \rightleftarrows c\,C + d\,D$

$K_p = \frac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$

$K_a = \frac{[H^+][A^-]}{[HA]}$

$K_b = \frac{[OH^-][HB^+]}{[B]}$

$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C

$= K_a \times K_b$

$pH = -\log[H^+]$, $pOH = -\log[OH^-]$

$14 = pH + pOH$

$pH = pK_a + \log\frac{[A^-]}{[HA]}$

$pK_a = -\log K_a$, $pK_b = -\log K_b$

Equilibrium Constants

K_c (molar concentrations)
K_p (gas pressures)
K_a (weak acid)
K_b (weak base)
K_w (water)

KINETICS

$\ln[A]_t - \ln[A]_0 = -kt$

$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$

$t_{1/2} = \frac{0.693}{k}$

k = rate constant
t = time
$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

$$n = \frac{m}{M}$$

$$\text{K} = {}^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

Molarity, M = moles of solute per liter of solution

$$A = abc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
M = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
a = molar absorptivity
b = path length
c = concentration

Gas constant, R = 8.314 J mol^{-1} K^{-1}
= 0.08206 L atm mol^{-1} K^{-1}
= 62.36 L torr mol^{-1} K^{-1}

1 atm = 760 mm Hg = 760 torr

STP = 273.15 K and 1.0 atm

Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$
$$= -RT \ln K$$
$$= -nFE^\circ$$

$$I = \frac{q}{t}$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

PERIODIC TABLE OF THE ELEMENTS

1												13	14	15	16	17	18
1 **H** 1.008	2																2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.30	3	4	5	6	7	8	9	10	11	12	13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.10	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.91	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.02	89 **†Ac** 227.03	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (271)	111 **Rg** (272)							

***Lanthanide Series**	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.5	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
†Actinide Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)